**IGCSE Chemistry notes sections 1 and 2 2019. Section 1 a: States of matter**

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| 1.1 understand the three states of matter in terms of the arrangement, movement and energy of the particles.  1.2 understand the interconversion between the three states of matter in terms of the names of the interconversions, how they are achieved, the changes in  arrangement, movement and energy of the particles. |

**The kinetic particle theory:**

|  |  |
| --- | --- |
| * all matter consists of tiny particles; these particles can be atoms, molecules or ions; * all particles move and, therefore, have kinetic energy; they vibrate or move position or both; * there are forces between the particles; | * when matter is heated its particles gain more kinetic energy and move faster or vibrate more; * when matter is heated it expands as its particles move further apart; * the temperature of a substance tells us the average amount of kinetic energy of all its particles . |

**The three states of matter**

We can use the kinetic theory to describe the three states of matter.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **State** | **Forces of attraction between particles** | Energy of particles | Arrangement of particles | **Movement/motion of particles** |
| **Solid** | Strong forces | Little energy | * Particles are held in fixed positions making an ordered arrangement or lattice * Particles are close together * When heated they move slightly further apart from each other | * Particles vibrate around a point * Particles cannot move from their positions * When heated they vibrate more and move further apart a little(=expand) |
| **Liquid** | Forces of attraction have less effect so that particles still stick together | Moderate energy | * Particles are in a random/irregular arrangement * Particles are still fairly close together * When heated they move slightly further apart from each other | * Particles can move past one another and move randomly * When heated they move faster and move further apart a little (=expand) |
| **Gas** | Forces of attraction have no effect – almost no forces holding them together | A lot of energy | * Random arrangement * Particles are very far apart/a lot of space between them/large distance between them * When heated they move even further apart | * Particles move randomly (= each particle in its own direction) and very fast * Particles move freely, constantly and in straight lines * Particles bounce off the walls of the containers causing pressure * When heated the particles move faster and move much further apart (=expand a lot) |

**Changes in state of matter**

The heating graph (energy is put in) below shows what happens to the state and temperature of a solid substance when it is heated.

**F**

time

Complete the table below

temperature

boiling

gas

|  |  |
| --- | --- |
| section | state(s) |
| A - B |  |
| B - C |  |
| C - D |  |
| D - E |  |
| E - F |  |

**D** **E**

Boiling point

Melting point

liquid

**B**

**C**

melting

solid

**A**

When a substance is heated its particles take in energy and use this energy to increase their movement and move further apart. PARTICLES OR MOLECULES DO NOT GET BIGGER (OR SMALLER!!) when they are heated (or cooled).

During the two sections in the graph above where the temperature remains, the substance is changing state; this is because when a substance changes state the heat energy supplied is used to overcome the effect of the forces of attraction instead of raising the temperature of the substance.

Fill in the gaps to describe what happens to the particles when a substance changes state by melting and boiling.

**liquid energy heated move positions vibrate boiling melting strong weak high bubbles low surface air evaporation**

|  |
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| Melting  When a solid is ………………………… its particles gain ………………….. and ……………. faster. Eventually they gain enough …………….  to move away from their ………………………. and begin to …………………….. around each other. The solid melts to form a ………………. . The  temperature at which a solid melts is the ……………………. point. The temperature at which a solid melts tells us how strongly its particles are held  together. Substances with high melting points have …………. forces of attraction between their particles. Substances with …… melting points have  weak forces between their particles. |

|  |
| --- |
| Boiling  When a liquid is heated the particles gain ………………. and …………….. around faster and faster. Some particles near the ………………….. of  the liquid have enough …………………. to escape into the ……….. and form vapour. This process is called ……………………………  Eventually a temperature is reached at which the particles are escaping from the liquid surface so rapidly that ……………………. of vapour (liquid  particles turned into gas particles) start to form inside the liquid. This process is called …………………………… |

Cooling graphs (energy is given out)

The cooling graph below shows what happens to the state and temperature of a gas when it is cooled.

temperature

condensing

During the cooling, the particles lose energy, come closer together and when they collide, they stick together to form droplets forming a liquid

If cooled further in the liquid state, they eventually move so slowly that start vibrating around a point.

gas

boiling point

freezing

liquid

melting point

solid

time

During condensing and freezing the particles lose energy as they need less energy because they slow down and come closer together again.

The forces between the particles have a much greater effect.

What are the changes of state called?

Complete the key to the right by writing the correct process next to each number.

1.

2.

3.

4.

5.

6.

**liquid**

**1 4**

**2 3**

**gas**

**solid**

**5**

**6**

**Evidence that matter is made from particles**

Dilution of solutions

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| 1.3 understand how the results of experiments involving the dilution of coloured solutions and diffusion of gases can be explained |

A blue copper sulphate solution can be made by dissolving copper sulphate crystals in water. The solution can then be used for a serial dilution.

A serial dilution is a series of stepwise dilutions, for instance a blue copper sulphate solution, usually by the same factor e.g. ten times; this can be done by taking for instance 10 cm3 of the copper sulphate solution and adding 90 cm3 of water to make 100 cm3 of a copper sulphate solution diluted ten times. The diluted copper sulphate solution has the same lighter blue colour throughout the new solution. This process can be repeated i.e. 10 cm3 is taken to which 90 cm3 of water is added to make up 100 cm3 again. Such a copper sulphate solution could be diluted further a thousand times or more and it will still show a blue colour.

The serial dilution of copper sulphate provides evidence that the initial copper sulphate crystal is made up of many small copper sulphate particles and that these particles spread out evenly when the solution is diluted by adding water.

Diffusion

Also the observations below can only be explained if we accept the idea that all matter is made up of particles that move, in the case of gases randomly.

* a balloon leaks air as it goes flat after a few days;
* making tea or coffee;
* two gases mix evenly with each other;
* the smell of food spreads from the kitchen;
* solids dissolve in liquids e.g. a sugar cube or a crystal of potassium permanganate: the solvent molecules cause the particles of the solid to separate and these dissolved solid particles then diffuse into the solvent.

In the boxes below show by using particles what happens when sugar has dissolved in water.

|  |  |
| --- | --- |
| Sugar cube in water before dissolving | Dissolved sugar cube |

Diffusion is the movement of particles with each particle moving in its own direction. This happens because particles with enough energy move from their positions i.e. particles in liquids and gases. Eventually particles from different substances will be mixed with each other.

In the boxes below draw particles to show how diffusion between involving 2 gases takes place.

|  |  |  |  |
| --- | --- | --- | --- |
|  |  | Diffusion |  |

**GAS A GAS B GAS MIXTURE OF A and B**

For a substance to be able to diffuse its particles must be able to move freely i.e. from their positions.

Although diffusion is fast in gases it still a slow process considering the speed at gas particles move.

**Diffusion is slow because gas and liquid particles all move in different directions. Diffusion is a random process (it happens as a result of the random movement of the particles)**

Diffusion is slower in liquids than in gases because liquid particles are not as free to move as gas particles. Gases and liquids can diffuse through solids but solids cannot diffuse through solids!!

Solutions

|  |
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| 1.4 know what is meant by the terms: solvent, solute, solution, saturated solution.  1.5 know what is meant by the term solubility in the units g per 100 g of solvent  1.6 understand how to plot and interpret solubility curves  *1.7 practical: investigate the solubility of a solid in water at a specific temperature* |

Terminology:

* Solute = a solid, liquid or gas that is dissolved in another substance.
* Solvent = substance, usually a liquid, that will dissolve another substance.
* Solution = a mixture made up from a solute and solvent.
* Saturated solution = a solution in which no more solute can be dissolved `at that temperature; the solution holds the maximum of solute.
* Solubility = is the **mass of solute** (in g) that can be dissolved in given **mass of solvent** i.e. 100g of a solvent (unit = g/100 solvent) at a given temperature, usually 25 0C. In the case of water as the solvent, the 100g of water is equal to 100 cm3 of water as water has a density of 1 g/cm3.

Solubility values allow us to compare the solubility of:

* different solutes, e.g. sugar and salt, at the same temperature
* a solute in different solvents e.g. in water and in alcohol
* a solute at different temperatures

Solubility curve

The relationship between the solubility of a solute in g/100g and the temperature of the solvent can be best represented by a solubility curve. As temperature is the independent variable it is set out on the horizontal axis (X-axis); solubility is the dependent variable and is therefore on the vertical or Y-axis.

The curves show how solubility of a solid changes with temperature. With most solutes the solubility increases with temperature but not all.

The solubility of most solids at a given temperature can be measured by experiment.

Exercises

1. Calculate the solubility (in g/100g) for the data collected for the following solutes at 25 0C.:
   1. 18g of sodium chloride dissolved in 50cm3 of water; assume density of water is 1g/cm3.
   2. 10g of copper sulphate dissolved in 25cm3 of water; assume density of water is 1g/cm3.
2. 11g of copper sulphate was dissolved in 25 cm3 of warm water. The first copper sulphate crystals appeared at 40 0C. Calculate the solubility of copper sulphate at 40 0C.
3. The table below shows the data collected from an experiment carried out to investigate the relationship between the solubility (in g/100g solvent) of potassium nitrate, sodium sulphate and copper sulphate and the temperature of the solvent.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Temperature (0C) | 0 | 20 | 40 | 60 | 80 | 100 |
| potassium nitrate | 13 | 33 | 65 | 106 | 167 | 240 |
| copper sulphate | 23 | 32 | 44 | 62 | 84 | 114 |
| sodium sulphate | 5 | 30 | 49 | 45 | 44 | 43 |

* 1. Use the data to draw solubility curves on graph paper for all three solutes.
  2. Why are there no solubility values after 1000C?
  3. What solute has the highest solubility at 00C? at 600C?
  4. For which solute does the temperature have the greatest effect on its solubility and how does the graph show that?
  5. What is the solubility at 70 0C for each solute; show how you have determined this value.
  6. What will you observe:
     1. When you add 50g of sodium sulphate to 100g of water at 60 0C?
     2. When you add 50g of copper sulphate to 50g of water at 100 0C
  7. What observations will you be able to make if a solution of 150g of potassium nitrate in 100g of water at 80 0C was cooled to 40 0C?
  8. 10g of copper sulphate was heated in 25 g of water and heated to 60 0C. At what temperature would crystals appear when the solution is allowed to cool?
  9. How much copper sulphate would you need to add to make a saturated solution in 100g of water at 50 0C?

1. The link below takes you to a solubility table in water on Wikipedia <https://en.wikipedia.org/wiki/Solubility_table>

Use the data in the table to answer the following questions:

* 1. How did the authors obtain all this data?
  2. Only consider solutes starting with a letter A, which solute has the highest solubility at 200C?
  3. Does the substance you have answered in b, also have the highest solubility at 1000C? Give a reason for your answer.
  4. How much ammonium chloride would you need to add to make a saturated solution at 800C?­­­

1. The solubility of sodium chlorate in water was measured at a number of different temperatures.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Temperature (0C) | 0 | 20 | 40 | 60 | 80 | 100 |
| Solubility (g/100g solvent) | 80 | 96 | 115 | 135 | 167 | 204 |

1. Use these figures to plot a solubility curve.
2. Use your graph to find the solubility at 500C
3. 110g of sodium chlorate was added to 100g of water and the mixture heated to about 700C. It was then left to cool with the thermometer in the solution. Use your graph to answer the following questions”
   1. At what temperature would crystals first appear in the solution?
   2. If the solution was cooled to 170C, work out the total mass of crystals formed.
4. A pair of students measured the solubility of a compound in water in the following way:

They placed 30.0 g of the solid in a small flask and added 20.0 cm3 of water. They heated the flask in a beaker of boiling water until all the solid had dissolved. Then they removed the flask from the hot water, placed a thermometer in the solution, and watched carefully for the first trace of crystals to appear. That shows that the solution had become saturated. They measured the temperature at that point.

Next they added some more water, and then repeated the experiment. They continued to do this until they had the set of results below.

Mass of solid used = 30.0 g

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Total volume of water used (cm3) | 20 | 30 | 40 | 60 | 80 | 100 |
| Temperature at which crystals appeared (0C) | 92.5 | 80.5 | 72.5 | 56.0 | 40.0 | 25.5 |

1. Calculate the solubility of the solid at each of the temperatures in grams of solid per 100g of water. Take the density of the water to be 1 g/cm3.
2. Plot a graph of solubility.
3. Use your graph to find the solubility of the solid at 500C.

**Section 1b: Elements, compounds and mixtures**

**Elements, compounds and mixtures**

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| 1.8 understand how to classify a substance as an element, compound or mixture.  1.9 understand that a pure substance has a fixed melting and boiling point, but that a mixture may melt or boil over a range of temperatures. |

All substances are made up three different types of particles two of which are called atoms and molecules (the third particle is called an ion – see ionic bonding). The smallest particle that can exist of a substance is called an atom. In many cases two or more atoms join together chemically to form a larger particle called a molecule.

Some molecules contain the same atoms, usually diatomic molecules such as H2, O2 and S8 whilst most molecules have different atoms joined together chemically, like H2O or C6H12O6.

We can divide up all substances into three groups by considering at the type of particle in them.

Elements

An element is a single substance that is made up of only **one** kind of atom and which cannot be split up into a simpler substance by any chemical process. More than 100 different elements have been identified, some of which have been made by man i.e. they are synthetic.

Compounds

Compounds are substances that are made when **two or more** elements combine together chemically; compounds are the products of a chemical change. A compound is made up of molecules or ions.

A compound has different properties from the elements in it! A compound also has more than two different types of particles/atoms/ions.

There are different types of compounds like organic and inorganic, acids, bases and salts, oxides, hydroxides and so on.

Mixtures

Mixtures are made up of different separate atoms or different molecules or a combination of both.

In the table below write 3 differences between a compound and a mixture.

|  |  |
| --- | --- |
| **compounds** | **mixtures** |
|  |  |

There are many ways which we can use to find out if a substance is pure or a mixture; you need to know two of these.

*Method 1: measuring the melting or boiling point of the substance*

All pure substances such as elements and compounds have fixed properties which means that they should always show these properties like boiling and melting points, conductivity, strength and so on.

Chemists have measured and know the melting points of (all) pure substances; you find most of these values in online databases. By testing a sample of a substance for its properties like melting and boiling point we can decide if the sample is a pure substance or not.

Example: if a clear liquid boils at 100 °C (or freezes/melts at 0° C) than the clear liquid must be water; if it boils at a temperature above 100 °C (or freezes/melts at a temperature below 0° C) than the water is not pure but is a mixture (a solution).

Pure substances have fixed melting or boiling points whilst mixtures boil or melt **over a range of temperatures** above the boiling point or below the melting point of the pure substances

Impurities usually increase the boiling point but decrease the melting point of a pure substance.

*Method 2: separating mixtures*

|  |
| --- |
| 1.10 Describe these experimental techniques for the separation of mixtures: simple distillation,  fractional distillation, filtration, crystallisation and paper chromatography |

Often the substances that we want to use or study are part of mixtures. If we want to study their properties or use them, they need to be pure (i.e. one substance only) as usually impurities change properties and make them less useful for our purpose.

You need to know the following separating techniques very well. This means that you should be able to describe how to carry out the separating technique (including being able to draw labelled diagrams of how to set up the apparatus) and know when to use it. For any given mixture you should be able to suggest the best separating technique.

Separating techniques use differences in properties between the substances that need separating. Some of these differences in properties are shown in the table below.

|  |  |  |
| --- | --- | --- |
| **Separating technique** | **How it works?** | **Diagram of set up** |
| Filtration | * Property used: **different size of particles** * How? Make mixture go through filter paper, component that does pass through is the residue. Other component makes up the filtrate | In your diagram below, label **residue** and **filtrate** |
| Crystallisation | * Property used: **different boiling points/decrease in solubility as temperature decreases** * How? Heat, substance with highest boiling point remains behind as crystals. As mixture cools solubility of solid decreases as crystalizes. |  |
| **Simple distillation** | * Property used: **different boiling points** * How? Boil mixture; lowest boiling point evaporates, collect steam and make it condense (2 stage process: evaporation and condensation) | In your diagram below label: **distillation flask, condenser, distillate** |

|  |  |  |
| --- | --- | --- |
| **fractional Distillation** | * Property used: **different but similar boiling points but** * How? Boil mixture; all components evaporate, collect steam, fractionating column causes condensation of highest boiling point, other component condenses in condenses ( 2 stage process: evaporation and condensation) | In your diagram below label: **distillation flask, condenser, distillate, fractionating column** |

The table below shows when each technique should be used.

|  |  |  |
| --- | --- | --- |
| **Separating technique** | **Type of mixture** | **example** |
| **Crystallisation** | Soluble solid and liquid – to obtain solid | Salt from salty water |
| **Filtration** | Insoluble solid and liquid | Sand and water |
| **Distillation** | Soluble solid and liquid – to obtain liquid/solvent | Pure water from salty water |
| **Fractional distillation** | Miscible liquids | Ethanol and water, crude oil |

**Fractional distillation**

#### Two liquids

Two liquids (e.g. ethanol and water) with different boiling points may be separated by fractional distillation. The liquid with the lower boiling point (e.g. ethanol) evaporates more easily and rises through a fractionating column. Any gas/vapour of the liquid with the higher boiling point condenses back down. Only the vapour of the liquid with the lower boiling point reaches the top of the column and diffuses into the condenser. From here it is condensed and separated from the liquid with the higher boiling point.

**Paper chromatography**

|  |
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| 1.11 understand how a chromatogram provides information about the composition of a mixture.  1.12 understand how to use the calculation of Rf values to identify the components of a mixture  1.13 practical: investigate paper chromatography using inks/food colourings |

How does paper chromatography work?

Place the steps in the correct order by writing numbers 1 to 7 in the right column. The first step should have number 1 next to it.

|  |  |
| --- | --- |
| Allow solvent to rise up the paper and dissolve the sample and to spread out the parts of the mixture on the chromatography paper. |  |
| Draw a line in pencil, called the **base line**, on chromatography paper (should be pencil as carbon does not dissolve in any solvent). |  |
| Place chromatography paper in suitable solvent. (solvent should not touch the sample on the paper). |  |
| Make a solution: If the sample is not a liquid then select a suitable solvent to make it into a solution: if insoluble in water, try ethanol or propanone. |  |
| Dissolve sample usually by crushing it in the solvent using a pestle and mortar. |  |
| Remove chromatography paper from beaker, allow the solvent to evaporate and interpret the chromatogram. |  |
| Place a small drop of the sample on base line on chromatography paper. |  |

Retention value Rf

In chromatography we can calculate a retention factor, or Rf. The retention factor can be defined as the distance traveled by the dye divided by the distance traveled by the solvent.

|  |  |
| --- | --- |
| Rf = | distance travelled by the dye |
| distance travelled by the solvent |

For example, if a compound travels 2.1 cm and the solvent front travels 2.8 cm, the Rf is 0.75.

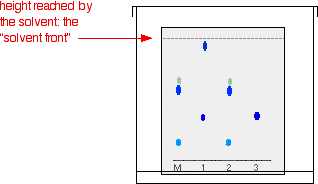
|  |
| --- |
|  |

Interpreting a chromatogram

Paper chromatography using propanone was used to find out if ink M from a pen was a pure substance or a mixture.

Study the chromatogram below and answer the questions.

(from <http://www.chemguide.co.uk/analysis/chromatography/paper.html> on23/4/2012).



1. Outline how the chromatogram was made.
2. What conclusions can you draw about substance M.
3. How is substance 3 different from the substances 1 and 2?
4. Suggest a reason why a lid has been placed on the beaker?
5. Why should the base line not be drawn in ink?
6. Why should the solvent not reach the base line at the start of the experiment?
7. Calculate the Rf value for
   1. Substance 3.
   2. The least soluble component in substance 2.
8. In another chromatography experiment the chromatogram showed a dot remaining on the base line. Suggest an explanation.

**Section 1c: atomic structure**

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| 1.14 know what is meant by the terms atom and molecule  1.15 know the structure of an atom in terms of the positions, relative masses and relative charges of sub-atomic particles.  1.16 know what is meant by the terms atomic number, mass number, isotopes and relative atomic mass (Ar)  1.17 be able to calculate the relative atomic mass of an element (Ar) from isotopic abundances |

**Atom and molecule**

What are the small particles that make up a substance? All substances are made up three different types of particles two of which are called atoms and molecules (the third particle is called an ion).

The smallest particle that can exist of a substance is called an atom. In many cases two or more atoms join together chemically to form a larger particle called a molecule. Some molecules contain the same atoms, usually diatomic molecules such as H2, O2 and S8, whilst most molecules have different atoms joined together chemically, like H2O or C6H12O6 and form a compound.

**Sub-atomic particles**

|  |  |  |  |
| --- | --- | --- | --- |
| **particle** | **relative mass** | **relative charge** | **where in atom?** |
| electron | 1/1840 |  |  |
| proton |  |  |  |
| neutron |  |  |  |

relative = as compared to …

The nucleus of each atom specie can be described by just two numbers:

1. the **atomic number, Z,** that is the number of protons in the nucleus of an atom;
2. the **mass number, A,** that is the number of protons and neutrons in the nucleus of an atom.

All this information about each atom can easily be shown in a short hand notation like the example below which shows the short hand notation of an iodine atom.

**mass number** 131

**I**

**atomic number** 53

The proton number is also equal to the charge on the nucleus and, as an atom is a neutral particle, to the number of electrons orbiting around the nucleus.

Using the short hand notation or your periodic table you should be able to find the number of protons, neutrons and electrons in each atom. Remember that in each atom:

* the number of protons is equal to the number of electrons
* the number of neutrons = the nucleon or mass number – proton number

Use the above information together with the Periodic Table to complete the table below.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **element** | **symbol** | **atomic number** | **number of protons** | **number of electrons** | **mass number** | **number of neutrons** |
| hydrogen |  | 1 |  |  |  | 0 |
| helium |  |  |  | 2 |  | 2 |
| carbon |  |  | 6 |  |  | 6 |
| nitrogen |  | 7 |  |  | 14 |  |
| oxygen |  | 8 |  |  | 16 |  |
| neon |  |  | 10 |  |  | 10 |
| sodium |  |  |  | 11 | 23 |  |
| magnesium |  | 12 |  |  |  | 12 |
| aluminium |  | 13 |  |  | 27 |  |
| sulphur |  |  |  | 16 | 32 |  |
| chlorine |  |  | 17 |  |  | 18 |

**Isotopes** (you should always use this word in the plural form)

**Isotopes** are atoms of the same element which have a different number of neutrons and therefore also a different mass number.

Example: Two common isotopes of chlorine are Cl - 35 and Cl – 37 (the numbers are their mass numbers);

1. 75% of any chlorine sample contains Cl - 35
2. and 25% of a chlorine sample contains Cl - 37.

Complete the table below which shows more examples of isotopes:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **isotope** | **atomic number** | **number of protons** | **number of electrons** | **mass number** | **number of neutrons** |
| Cl – 35 |  |  |  |  |  |
| Cl - 37 |  |  |  |  |  |
| H - 1 |  |  |  |  |  |
| H - 2 |  |  |  |  |  |
| H - 3 |  |  |  |  |  |
| C- 12 |  |  |  |  |  |
| C - 14 |  |  |  |  |  |

**Relative atomic mass, Ar**

The relative atomic mass of an element is the weighted average mass of all isotopes of an element compared to the 1/12th of the mass of a C-12 isotope.

The relative atomic mass can also calculated as shown below:

|  |  |
| --- | --- |
| **Ar =** | (relative isotopic mass 1 x %) + (relative isotopic mass 2 x %) |
| 100 |

It is the relative atomic mass that is found in the Periodic Table.

Example: the relative atomic mass of chlorine is 35.5 because

|  |  |  |
| --- | --- | --- |
| **Ar =** | (35 x 75) + (37 x 25) | = 35.5 |
| 100 |

or

**Ar =** (35 x 0.75) + (37 x 0.25) = 35.5.

Complete the following table:

|  |  |  |  |
| --- | --- | --- | --- |
| **element** | **isotopes** | **abundance** | **relative atomic mass** |
| Cu | Cu-63  Cu-65 | 69 %  31 % |  |
| Ne | Ne-20  Ne-22 | 90%  10% |  |
| B | B-10  B-11 | 20%  80% |  |
| Br | Br-79  Br-81 | 50.5%  49.5% |  |
| Ag | Ag-107  Ag-109 | 48% |  |
| Mg | Mg-24  Mg-25  Mg-26 | 79%  10%  11% |  |

(more on <http://www.ausetute.com.au/atomicmass.html>)

**Section 1d: the Periodic Table**

|  |
| --- |
| 1.18 understand how elements are arranged in the Periodic Table:   * in order in order of atomic number * in groups and periods   1.19 understand how to deduce the electronic configurations of the first 20 elements from their positions in  the Periodic Table  1.22 understand how the electronic configuration of a main group element is related to its position in the  Periodic Table  1.23 understand why elements in the same group of the Periodic Table have similar chemical properties  1.24 understand why the noble gases (Group 0) do not readily react |

**The Periodic Table**

The periodic table is a method of classifying elements into groups and periods.

1. Modern periodic table: elements are arranged in order of increasing atomic number.
2. **group** = vertical column of elements; within a group there are distinct chemical similarities but also gradual changes in properties as we will see in groups 1 and 7.
3. **period** = horizontal row of elements from an alkali metal to noble gas.

**Electron configuration of elements**

The electrons in an atom orbit the nucleus in a series of **electron** **shells** (energy levels) which are at an increasing distance away from the nucleus.

The electron configuration in an atom tells how the electrons are arranged in the different energy shells. You should be able to draw the electronic structure of the first 20 elements in the Periodic Table.

When determining the electron configuration of an atom, we need to follow a few simple rules:

* the first shell (the lowest energy level) is always filled first, then the second and then the third; you start putting the electrons on the shells from the nucleus outwards i.e. fill up the first shell first and then the second and so on.
* each electron shell can only hold a certain amount of electrons: in order of increasing distance from the nucleus, these are:

|  |  |
| --- | --- |
| *first shell* | *up to 2 electrons* |
| *second shell* | *up to 8 electrons* |
| *third shell* | *up to 8 electrons* |

Using examples from your textbook and the worksheet “Electron arrangement” draw the electronic configurations of helium, fluorine, sodium, aluminium and potassium using energy shells.

**Electron configuration and the Periodic Table**

We can also show the electronic structure of an atom in a shorter way (without drawing the shells) as shown by the ones for lithium and argon in the section of the Periodic Table.

Complete the short electronic configuration for all other elements in the diagram below.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| group | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 0 |
| period 1 | H |  |  |  |  |  |  | He |
| period 2 | Li  **2,1** | Be | B | C | N | O | F | Ne |
| period 3 | Na | Mg | Al | Si | P | S | Cl | Ar  **2,8,8** |
| period 4 | K | Ca |  |  |  |  |  |  |

The diagram above shows the link between the position of an element in the periodic table and its electronic configuration as summarized by the table below.

|  |  |
| --- | --- |
| **Group number and number of electrons in outer shell are the same** | The *number of outer shell electrons* or *valency electrons* determines the **group** number. The zero in Group refers to the shell beyond the outer one, which does not have any electrons in it yet (it would be better to call it group 8); |
| **Period number and number of shells are the same** | The number of the *shell* which is being filled determines the **period** to which an element belongs; |
| **Noble gas electron configuration** | Using the completed diagram above you will notice that all the elements in **Group 0**, the **noble gases**, have filled outer shells. This is a very stable electronic structure which makes these elements **unreactive** or inert.    examples: He : 2 Ne : 2, 8 Ar : 2, 8, 8 |

To summarise: An element’s:

* group number is the same as the number of valence electrons in its outer most shell.

1. period number is the same as the number of energy shells that have electrons on them or the number of the energy shell being filled.

**Elements in same group have similar chemical properties**

**The chemical properties of an element depend on the number of electrons in its most outer energy shell**.

As all elements in the same group have the same number of valence electrons all elements in the same group will have similar chemical properties. For instance, all elements in group 1 have 1 electron in their last shell and because of that are all very reactive.

**Noble gases are unreactive**

The elements in the last group, group 0, are inert gases. They are unreactive because they have a full outer shell (2 electrons in the case of helium and 8 electrons in the other noble gases) unlike the elements in the other group.

**Metals and non-metals**

|  |
| --- |
| 1.21 identify an element as a metal or a non-metal according to its position in the Periodic Table.  1.20 understand how to use electrical conductivity and the acid-base character of oxides to classify  elements as metals or non-metals |

All elements in the periodic table are divided up into metals and non-metals. The metals are found to the left and the middle whilst the non-metals on the right.

The table below summarizes the differences in properties between metals and non-metals most of which you already know.

|  |  |
| --- | --- |
| **Metals** | **Non-metals** |
| * Usually have high melting and boiling points; they are all solids except mercury; * Malleable * Good conductors of heat and electricity; * Usually shiny (most are also silvery); * Ductile * Their oxides are basic i.e. they have a high pH in solution and/or can neutralize and acid. | Variable melting and boiling points but usually lower than in metals; most non-metals are either solids or gases;  Brittle;  Poor conductors of heat and electricity (except graphite);  Non-metal oxides are acidic (low pH in solution) |

The 2 properties focused on in this course are in the statements above.

1. The reason why metals are very good conductors of electricity and non-metals are poor conductors is because the atoms in metals are bonded differently to the atoms in non-metals; more in this in sections 1g and 1h.
2. If an oxide (e.g. carbon dioxide) has a pH lower than 7 then the element that has combined with oxygen to form the oxide is a non-metal. If the oxide, e.g. copper oxide, is a basic oxide with a pH above pH 7 that can neutralize an acid, the element is a metal.

**Section 1e: chemical formulae, equations and calculations**

|  |
| --- |
| 1.25 write word equations and balanced chemical equations (including state symbols) for reactions  studied in this specification and for unfamiliar reactions where suitable information is provided |

**Word equations**

|  |
| --- |
| Write word equations for the following chemical equations. |
| a. CaCO3 + 2HCl → CaCl2 + H2O + CO2 |
| b. H2SO4 + 2KOH → K2SO4  + 2H2O |
| c. Ca + 2H2O → Ca(OH)2 + H2 |
| d. Fe2O3 + 3H2 → 2Fe + 3H2O |
| e. 2CO + O2 → 2CO2 |
| f. 4Fe + 3O2 → 2Fe2O3 |
| g. 2Al + 3I2 → 2AlI3 |
| h. C2H4 + 3O2 → 2H2O + 2CO2 |
| i. 2CuSO4 + 4KI → 2K2SO4  + 2CuI + I2 |
| j. 2Ag2CO3 → 4Ag + 2CO2 + O2 |

**Balancing chemical equations**

Important points to remember!!!!

* **The number of atoms on both sides must be the same!!!!!**
* **You can only write numbers in front of the formulae**
* **You cannot change any formulae of the compounds involved – a common error**

|  |  |
| --- | --- |
| 1. Check if the following symbol equations are balanced or not. If they are not you will need to balance them in your book. | |
| * 1. C + O2 ⎯→ CO2 | h. Ca + O2 ⎯→ CaO |
| * 1. CO + O2 ⎯→ CO2 | 1. CH4 + O2 ⎯→ H2O + CO2 |
| * 1. CuO + H2 ⎯→ Cu + H2O | j. KBr + Cl2 ⎯→ KCl + Br2 |
| * 1. Zn + CuSO4 ⎯→ ZnSO4  + Cu | k. H2 + Cl2 ⎯→ HCl |
| * 1. HCl + MgO ⎯→ MgCl2 + H2O | l. N2 + H2 ⎯→ NH3 |
| * 1. H2 + O2  ⎯→ H2O | m. H2 + NO ⎯→ H2O + N2 |
| * 1. Al + Cl2 ⎯→ AlCl3 | n. Fe + O2 ⎯→ Fe2O3 |

|  |
| --- |
| 2. Write balanced equations from the information given below (wherever you can include state  symbols)  **(s) = solid (l) = liquid (g) = gas (aq) = dissolved in water** |
| 1. Copper (II) oxide reacts with hydrogen to form water and copper. |
| 1. Sulphur dioxide reacts with oxygen to form sulphur trioxide. |
| 1. Carbon monoxide reacts with oxygen to form carbon dioxide. |
| 1. Magnesium reacts with hydrochloric acid to form hydrogen and magnesium chloride. |
| 1. Calcium hydroxide solution reacts with sulphuric acid to give calcium sulphate and water. |
| 1. Zinc reacts with steam to form zinc oxide and hydrogen |
| 1. Magnesium carbonate decomposes to form magnesium oxide and carbon dioxide. |
| 1. Carbon and carbon dioxide react to form carbon monoxide. |
| 1. Calcium reacts with water to form hydrogen and a solution of calcium hydroxide. |
| 1. Sulphur burns in oxygen to form sulphur dioxide. |
| 1. Aluminium and bromine combine to from aluminium bromide |
| 1. Methane gas burns in oxygen to form carbon dioxide and steam. |
| 1. Sodium reacting with water (you should remember the products) |
| 1. Rubidium reacting with water |
| 1. Chlorine reacting with potassium iodide |

**Relative formula mass, Mr**

|  |
| --- |
| 1.26 calculate the relative formula masses (including relative molecular masses (*Mr*) from relative atomic masses (*Ar*) |

The relative formula mass is the relative mass of compounds and is calculated using the relative atomic masses of the elements in the compound multiplied by the number of particles of that element as shown below:

|  |
| --- |
| * relative molecular mass of H2SO4 is 98 because (2 x 1) for H + 32 for S + (4 x 16) for O = 98 |
| * relative molecular mass of H2O is 18 because (2 x 1) for H + (1 x 16) for O = 18 |
| * relative formula mass of NaOH is 40 because 23 + 16 + 1 = 40 |
| * relative molecular mass of Al2(SO4)3 is because (2 x 27 ) + (3 x 32) + (12 x 16) = 342 |

Find the relative molecular mass for each of the following. SHOW HOW YOU DID IT!!!

1. HCl b. MgCl2 c. LiOH d. CH4 e. CaCO3 f. Mg(NO3)2 g. CaSO4

h. NH4OH i. (NH4)2 SO4 j. H2SO4 k. O2 l. N2 m. CuSO4.5H2O

**The mole**

|  |
| --- |
| 1.27 know that the mole (mol) is the unit for the amount of a substance.  1.28 understand how to carry out calculations involving amount of substance, relative atomic mass (Ar)  and relative formula mass (Mr)  1.29 calculate reacting masses using experimental data and chemical equations |

Because atoms and molecules are so small, chemists measure or work with them in large numbers.

Chemists measure amounts of substance in moles.

The mole is just a name used to indicate a certain number, just like a pair, a dozen, a trio, and that number is 6.02 x 1023  which also called Avogadro's number.

Therefore a mole always equals 6.02 x 1023  so we have a mole of atoms or molecules or ions or electrons or ….

How can we take 1 mole of a substance or any other number of moles?

In the case of solids we can use the molar mass.

**Molar mass**

The molar mass of a substance is the mass of 1 mole of particles of that substance or the mass of

6.02 x 1023 particles of that substance.To make it easy, the molar mass happens to be the relative atomic mass (*Ar*) or relative formula mass (*Mr*) but expressed in grams.

**Calculating moles using molar mass**

Once we know the mass of a substance we can now also calculate how many moles of the substance we have by using the following VERY IMPORTANT FORMULA!!!

|  |  |
| --- | --- |
| number of moles = | mass of substance |
| *molar mass* |

Find the number of moles of the following – show your working.

a. 73g of HCl b. 160g of bromine **molecules** c. 10g of sodium hydroxide

d. 85g of NH3 e. 30g of magnesium sulphate f. 0.8g of CH4

g. 196g of H2SO4  h. 31g of P4 i. 0.56g of CaO

Find the mass of the following amounts of substance– show your working:

1. 2 moles of iron b. 0.5 mole of water c. 100 moles of HNO3

d. 0.005 mole of CuSO4  e. 0.1 mole of Ca(OH)2  f. 20 moles of chlorine **molecules**

g. 0.25 moles of CaCO3 i. 0.1 moles of S8

**Calculations of reacting masses**

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| ***Example*** How much iron oxide is needed to produce 28 g of iron?  *Step 1: write balanced chemical equation:*  Fe2O3(s) + 3CO(g) → 2Fe (s) + 3CO3 (g)  S*tep 2: convert data into moles:*   |  |  |  |  |  | | --- | --- | --- | --- | --- | | number of moles = | mass |  | 28g |  | |  | = |  | = 0.5 mole of Fe | | molar mass |  | 56g |  |   S*tep 3: find answer in moles using the balanced chemical equation:*  2 moles of iron needs 1 mole of iron oxide  ∴ 0.5 moles of iron needs 0.25 moles of iron oxide  *Step 4: convert answer into unit asked for:*  unit is mass → mass = number of moles x Mr = 0.25 x 160g = 40 g |

Exercises

|  |
| --- |
| 1. How much zinc oxide will be produced when 13g of zinc are heated in an excess of oxygen? 2. How much calcium oxide will be produced when 10g of calcium carbonate are decomposed by heat? 3. 6.4 g of sulphur were burned completely in air. What mass of sulphur dioxide was formed? 4. How much carbon dioxide is produced when 20 tonnes of carbon are burned in excess of oxygen? 5. How much oxygen is needed to react with 12 tonnes of iron to form iron (III) oxide? 6. A chemist is investigating a mixture of substances to make an improved type of glass. She wants   the finished glass sample to contain 14.0 g of calcium oxide. She plans to add calcium carbonate to  the mixture before it is melted. Calcium carbonate undergoes thermal decomposition according to  the equation  CaCO3(s) → CaO (s) + CO2 (g)  Calculate the minimum number of moles of calcium carbonate which the chemist should add to the  mixture in order to ensure that the final glass contains 14.0 g of calcium oxide.  7. Calculate the mass of copper required to produce 160g of copper oxide when heated in excess  oxygen  2Cu (s) + O2(g) → 2CuO (s)   1. Zinc oxide reacts with carbon to form zinc and carbon monoxide. How much zinc oxide do you   need to make 130 tonnes of zinc?  ZnO (s) + C(s) → Zn (s) + CO (g)   1. In the thermit reaction, aluminium is used to obtain pure iron. How much aluminium is needed to obtain 28 tonnes of iron?   Fe2O3(s) + 2Al (s) → 2Fe (s) + Al2O3 (s)   1. What mass of chlorine would be needed to make 73 tonnes of hydrogen chloride?   H2 (g) + Cl2 (g) → 2HCl(g)   1. Calculate the mass of nitric acid needed to make 400 tonnes of ammonium nitrate   HNO3 + NH3 → NH4NO3   1. What mass of aluminium would be produced from 204 tonnes of aluminium oxide given the reaction?   2Al2O3 (s) → 4Al (s) + 3O2 (g)   1. 2.67g of aluminium chloride was dissolved in water and silver nitrate was added to give silver chloride   AlCl3  + 3AgNO3  → Al(NO3) + 3AgCl  What mass of silver chloride was produced?   1. Calcium hydroxide is manufactured by heating calcium carbonate strongly to produce calcium   oxide, and then adding a controlled amount of water to produce calcium hydroxide.  CaCO3(s) → CaO (s) + CO2 (g) CaO (s) + H2O (l) → Ca(OH)2 (s)  a. What mass of calcium oxide would be produced from 1 tonne of calcium carbonate?  b. What mass of water would you need to add to that calcium oxide?  c. What mass of calcium hydroxide would eventually be produced?   1. Chromium is manufactured by heating a mixture of chromium (III) oxide with aluminium powder.     Cr2O3(s) + 2Al (s) → 2Cr (s) + Al2O3 (s)  a. Calculate the amount of aluminium needed to react with 1 tonne of chromium (III) oxide.  b. Calculate the amount of chromium produced from 1 tonne of chromium (III) oxide |

**Percentage yield**

|  |
| --- |
| 1.30 Calculate percentage yield |

The amount of a product that you could obtain according to calculations using the ratio’s from the equation is called the theoretical yield.

However, quite often, especially with organic reactions, the calculated amount of product is not obtained when the reaction is carried out.

Percentage yield indicates the percentage of the theoretical yield that has been obtained and can be calculated using the following formula:

percentage yield = obtained yield/theoretical yield x 100

Example: synthesis of ammonia has a 90 %yield! This means that 90% of what we should get according to the calculations is obtained when the reaction is carried out.

Exercises:

|  |
| --- |
| 1. Zinc and chlorine react directly to give zinc chloride  Zn (s) + Cl2 (g) → ZnCl2 (s)   1. If you begin with 1.00 mole of zinc and excess Cl2, what is the theoretical yield of ZnCl2 in grams? 2. If you make 115 g of ZnCl2 what is the percent yield of the chloride? |
| 1. Nickel tetracarbonyl, Ni(CO)4, is made by direct reaction of nickel metal and CO gas.   Ni (s) + 4CO (g) → Ni(CO)4 (s)   * 1. If you begin with 5.00 g of Ni and an excess of CO, what is the theoretical yield of Ni(CO)4 ?   2. How many grams of CO are required to use up all of the nickel?   3. If you make only 9.67 g of Ni(CO)4, what is the percent yield of the compound? |
| 1. The predicted yield of a reaction was worked out and found to be 12.5g. However, when the reaction was carried out, the actual yield was only 8.25g. Calculate the percentage yield. |

**Finding chemical formulae: experimental methods**

|  |
| --- |
| * 1. understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallisation   1.32 know what is meant by the terms empirical formula and molecular formula  1.33 calculate empirical and molecular formulae from experimental data  1.36 practical: know how to determine the formula of a metal oxide by combustion (e.g. magnesium  oxide) or by reduction (e.g. copper(II) oxide) |

Chemical formulae of chemical compounds can be obtained by conducting experiments with the compounds to determine how much there is (in grams or in percentage) of each element in a compound.

There are two common experimental methods which allow us to obtain this information. Which method is used depends on the type of compound. We will restrict ourselves to two types of compounds only and therefore two types of experimental methods.

Metal oxides e.g. magnesium oxide:

The method involves heating strongly a known mass of magnesium. During heating the magnesium reacts with oxygen in the air and its mass increases as it changes into magnesium oxide a shown by the equation below.

Mg (s) + O2 (g) ⎯→ MgO (s)

We need to measure the increase in mass as it equals the amount of oxygen which has reacted with the magnesium. We now have both masses necessary to calculate the empirical formula. The method for calculating the empirical formula will be shown below.

Complete the table below

|  |  |  |
| --- | --- | --- |
| mass of magnesium (g) | mass of oxygen (g) | mass of magnesium oxide(g) |
| 0.025 |  | 0.042 |
|  | 0.004 | 0.010 |
| 0.160 | 0.109 |  |

Hydrated salts e.g. CuSO4.5H2O

When a hydrated salt, such CuSO4.5H2O, is heated it decomposes and the water of crystallization from the compound evaporates. The mass of the salt decreases allowing us to calculate the mass of the dry anhydrous salt, CuSO4, the mass of the water of crystallization and the formula of the hydrated salt as shown below.

CuSO4.5H2O → CuSO4 + 5 H2O

The compound is heated until its mass does not decrease any further.

**Empirical formula**

The empirical formula shows the **simplest whole number molar ratio of the atoms/ions in a molecule/formula unit** and can be calculated from experimental results, i.e. masses, using moles. The empirical formula can also be used to indicate the ratio of salt particle to water molecules in a hydrated salt.

|  |
| --- |
| For each of the molecular formula below, find the empirical formula.  a) H2O b) H2O2 c) C6H6 d) C2H6 e) C2H4 f) C4H8O2  g) NH3 h) Fe2O3 I) Al2Cl6 |

## Example 1: An experiment shows that 64g of compound contains 48g carbon and 16g hydrogen. Calculate its empirical formula.

|  |  |  |
| --- | --- | --- |
| elements | carbon | hydrogen |
| *mass (in g)* | 48 | 16 |
| molar mass | 12 | 1 |
| number of moles | 4 | 16 |
| most simple molar ratio | 4/4 = 1 | 16/4 = 4 |
| empirical formula | CH4 | |

The above mass data can also be expressed in percentage.

|  |  |  |
| --- | --- | --- |
| elements | carbon | hydrogen |
| *mass (in %)* | 75 | 25 |
| molar mass | 12 | 1 |
| number of moles | 75/12 = 6.25 | 25/1 = 25 |
| most simple molar ratio | 6.25/6.25 = 1 | 25/6.25 = 4 |
| empirical formula | CH4 | |

**Example 2**: magnesium oxide using the data from the table on page 4.

|  |  |  |
| --- | --- | --- |
| elements | Mg | O |
| *mass (in g or %)* | 0.025 | 0.017 |
| molar mass | 24 | 16 |
| number of moles | 0.025/24 = 0.001 | 0.017/16 = 0.001 |
| most simple molar ratio | 1 | 1 |
| empirical formula | MgO | |

Exercises

|  |  |  |
| --- | --- | --- |
| 1. An experiment shows that 32g of sulphur combine with 32g of oxygen to form the compound sulphur dioxide. What is the empirical formula? 2. To form iron (II) sulphide, 56g of iron combines with 32g of sulphur. Find its empirical formula. 3. Calculate the empirical formulas of the compounds whose mass ratios are:  |  |  | | --- | --- | | * 1. 60% magnesium; 40 % oxygen   2. 59% sodium; 41% sulphur   3. 87.5% silicon: 12.5 % hydrogen | * 1. 79.8% copper: 20.2% sulphur   2. 42.8% carbon: 57.2% oxygen |  1. Calculate the empirical formulas of the compounds whose mass ratios are:    1. 52.1% carbon; 13.0% hydrogen; 34.9 % oxygen    2. 66.3% chlorine; 26.2 % nitrogen: 7.5% hydrogen    3. 62.2% iron: 35.6% oxygen: 2.2.% hydrogen    4. 40.0% calcium: 12.0% carbon: 48.0% oxygen    5. 1.46% hydrogen: 51.82% chlorine: 46.72% oxygen    6. 0.99% hydrogen: 35.4% chlorine: 63.61% oxygen 2. Compound X is a hydrocarbon: it contains only carbon and hydrogen atoms. 0.84g of X was   completely burned in air. This produced 2.64g of CO2 and 1.08g of H2O. Find the empirical formula.   1. A volatile compound has the composition by mass: arsenic 62.5 %, carbon 30.0 %, hydrogen 7.5   %. Calculate the empirical formula.  7. 4.2 g of cerium reacted with oxygen to form 5.16 g of an oxide of cerium. Complete the following to  determine the formula of this oxide.   1. Number of moles of cerium atoms used 2. Mass of oxygen that reacted 3. Number of moles of oxygen atoms in oxide 4. Ratio by moles of cerium atoms to oxygen atoms   Formula of this oxide of cerium is ……… |

**Molecular formula**

A molecular formula **shows the actual number of atoms/ions in a compound and is a simple multiple of the empirical formula.**

There are many possible molecular formula from one particular empirical formula. However, if the molecular mass of the substance is known, the molecular formula can be calculated using the method shown below.

**Example:**  Find the molecular formula of a compound which has CH2 as its empirical formula and has a molar mass of molecular formula of 84 g.

|  |
| --- |
| Method:  Calculate the mass of the empirical formula: CH2 = (12g x 1) + (1g x 2) = 14 g |
| Divide molar mass by empirical mass to find the multiple, n:  n = molar mass/empirical mass = 84g /14g = 6 |
| Multiply empirical formula by the multiple to find the molecular formula:  molecular formula= n x empirical formula = (CH2) x 6 = C6H12 |

Exercises

|  |
| --- |
| 1. A compound has the empirical formula CH2 and a formula mass of 28. What is the molecular formula? 2. An oxide of phosphorus has a formula mass of 220 and is 56.4% phosphorus. Find its molecular formula. 3. A gas has the empirical formula CH4. 0.16g of the gas occupies a volume of 240 cm3 at rtp. What is its molecular formula? 4. Octane is a member of the alkane family. Its percentage composition is 84.2% carbon and 15.8% hydrogen. Its formula mass is 114. What is its molecular formula? 5. 1,2-dibromoethane reacts with sodium hydroxide to form a compound that has the composition by mass: carbon, 38.2 %; hydrogen, 9.8 %; oxygen, 51.6%.   The relative molecular mass of the compound is 62. What is the molecular formula?  6. The Mr of oxalic acid is 90 and its composition by mass is: carbon: 26.7 % oxygen 71.1 %  hydrogen 2.2 %. What is the molecular formula of the acid?  7. An organic compound contained 66.7% carbon, 11.1% hydrogen and 22.2% oxygen. Its relative  formula mass is 72. Find its empirical and molecular formula. |

**Hydrated salts:** When hydrated magnesium sulphate, MgSO4.nH2O, was heated until there was no further mass decrease 4.58 g of anhydrous magnesium was left behind and 4.86g of water was released.

|  |  |  |
| --- | --- | --- |
| compound | MgSO4 | H2O |
| *mass (in g or %)* | 4.58 | 4.86 |
| molar mass | 120 | 18 |
| number of moles | 4.58/120 = 0.039 | 4.86/18 = 0.27 |
| most simple molar ratio | 0.039/0.039 = 1 | 0.27/0.039 = 6.93 |
| empirical formula | MgSO4.7H2O | |

Exercises

|  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 1. 416 g anhydrous barium chloride were obtained when 488g of the hydrated salt were heated.  Calculate *n* in the formula BaCl2.*n*H2O  2. A sample of magnesium sulphate crystals weighing 0.942 g was heated to drive off the water of  crystallization. When it reached constant mass, the mass of the residue was 0. 461g. Calculate  the empirical formula of the hydrate   1. A sample of calcium sulphate crystals weighing 1.124 g was heated to drive off the water of crystallization. When it reached constant mass, the mass of the residue was 0.889 g. Calculate the empirical formula of the hydrate. 2. The chemical name for **gypsum** is calcium-2-sulphate. A student followed these directions for an experiment with **gypsum**. Weigh a crucible empty. Weigh accurately about 30 g of gypsum into the crucible. Heat the crucible strongly for 10 minutes. Then allow it to cool, and reweigh. Heat the crucible again; cool and reweigh. The results she obtained are listed below.  |  |  | | --- | --- | |  | Mass (in g) | | Mass of crucible | 12.305 | | Mass of crucible + gypsum | 46.705 | | Mass of crucible + gypsum after heating | 41.305 | | Mass of crucible + gypsum after heating | 41.305 | |  |  |  1. Find the percentage loss in mass when gypsum is heated. 2. Tests on the product of heating show the presence of calcium and sulphate. The tests prove that only water has been lost. If the formula of gypsum is CaSO4.2H2O, find the formula of the product by means of the following steps    * Mr of gypsum =    * mass of 1 mole of gypsum = g    * percentage of this mass lost as water = %    * mass of water lost from 1 mole of gypsum = g    * Mr of water =    * number of moles of water lost from 1 mole of gypsum =    * the formula of the product is …   5. 7.50 g of hydrated copper sulphate crystals are gently heated until the loss of water of crystallization  appears to be complete. After cooling they have a mass of 5.35g. They are then reheated and  cooled and their mass is found to be 4.80g. Further heating does not change this mass.   * 1. In this experiment, how would you judge if the reaction was complete?   2. Do you think the reaction is finished after the first heating? Explain your answer.   3. Do you think the reaction is finished after the second heating? Explain your answer.   4. Work out the number of mol H2O and the number of mol CuSO4 originally present in the crystals.   5. Work out the number of mol H2O which combine with 1 mol CuSO4 in the crystals.   Write the formula for hydrated copper sulphate and the equation for its thermal decomposition to anhydrous copper sulphate. |

**Calculations using volumes and solution**

|  |
| --- |
| **1.34C understand how to carry out calculations involving amount of substance, volume and concentration (in mol/dm3) of solution** |

Many reactions take place with the reactants dissolved in water. Of each solution it is important that we know or can calculate its concentration.

**Concentration of a solution = the amount of substance dissolved in a 1dm3 of solution (in mol/dm3).**

**Concentration of a solution can also be measured in g/dm3.**

We can calculate the concentration by using the following formula:

|  |  |
| --- | --- |
| concentration = | number of moles |
| volume in dm3 |

**Exercises**

1. How many moles of solute are there in:

|  |
| --- |
| a. 500 cm3 of solution, concentration 2 mol/dm3 |
| b. 2 litres of solution, concentration 0.5 mol/dm3 |
| c. 20 cm3 of solution, concentration 2 mol/dm3 |

2. What is the concentration of a solution containing

|  |
| --- |
| a. 4 moles in 2 dm3 of solution |
| b. 0.5 moles in 0.1 dm3 of solution |
| c. 3 moles in 200 cm3 of solution |

3. What volume of

|  |
| --- |
| a. a 4 mol/dm3 solution contains 2 moles? |
| b. a 6 mol/dm3 solution contains 0.03 moles? |

4. The formula mass of sodium hydroxide is 40. How many grams of sodium hydroxide are there in

|  |
| --- |
| a. 500 cm3 of solution, concentration 1 mol/dm3 |
| b. 25 cm3 of solution, concentration 0.5 mol/dm3 |

5. What is the concentration of:

a. a sodium carbonate solution containing 53g of the salt Na2CO3 in 1 litre?

b. a copper sulphate solution containing 62.5g of CuSO4.5H2O in 1 litre?

6. What mass of silver nitrate, AgNO3, would you need to make 50 cm3 of a 0.2 mol/dm3 solution?

Calculating reacting amounts using concentrations

|  |
| --- |
| How do we calculate a volume or a concentration of a reacting solution? |
| Example: What volume of sulphuric acid, of a concentration of 1 mol/dm3, is needed to react completely with 500 cm3 of sodium hydroxide with a concentration of 2 mol/dm3? |
| *step 1: write balanced equation:*  2NaOH (aq) + H2SO4(aq) → 2H2O (l) + Na2SO4(aq) |
| *step 2: convert data into moles:*  number of moles = concentration x volume  = 2 mol dm3 x 0.5 dm3 (=500 ml) = 1 mole of NaOH |
| *step 3: find answer in equation:*  2 moles of sodium hydroxide need 1 mole of sulphuric acid  1 mole of sodium hydroxide needs 0.5 moles of sulphuric acid |
| *step 4: convert answer into unit asked for*   |  |  |  |  |  | | --- | --- | --- | --- | --- | | volume = | number of moles of sulphuric acid | = | 0.5 moles | = 0.5 dm3 | | concentration of sulphuric acid | 1 mol/dm3 | |

Exercises

1. 30 cm3 of 0.2 mol/dm3 hydrochloric acid just neutralized 20 cm3 of a sodium hydroxide solution. What is the concentration of the sodium hydroxide solution used?
2. 40 cm3 of 0.1 mol/dm3 sulphuric acid just neutralized 20 cm3 of a potassium hydroxide solution. What is the concentration of the potassium hydroxide solution?
3. 25.0 cm3 of 0.100 mol/dm3 sodium hydroxide solution required 23.5 cm3 of dilute hydrochloric acid for neutralization. Calculate the concentration of the hydrochloric acid.
4. A sample of hard water contained 0.002 mol/dm3 of calcium hydrogen carbonate, Ca(HCO3)2.

When this is heated, it decomposes to make calcium carbonate. Calculate the mass of calcium

carbonate which could be formed when 100 dm3 of the hard water is heated.

5. 25 cm3 of sodium hydroxide solution of unknown concentration was titrated with dilute sulphuric

acid of concentration 0.050 mol/dm3. 20.0 cm3 of the acid was required to neutralize the alkali.

Find the concentration of the alkali.

Additional exercises

1. Complete the table below.

|  |  |  |
| --- | --- | --- |
| **concentration (in mol/dm3)** | **volume** | **number of moles** |
| 0.5 | 1000 cm3 |  |
|  | 1000 cm3 | 2 |
| 2 | 2000 cm3 |  |
| 2.5 |  | 1 |
| 0.01 |  | 10 |
|  | 200 cm3 | 0.02 |
| 2 | 5 dm3 |  |

2. How many grams of solute are needed to make the following solutions.

|  |  |  |  |
| --- | --- | --- | --- |
| **substance** | **concentration(mol/dm3)** | **volume** | **mass of solute(g)** |
| **CuSO4** | 2 | 500 cm3 |  |
| **NaOH** | 0.5 | 1 dm3 |  |
| **HCl** | 1 | 100 cm3 |  |
| **AgNO3** | 2.5 | 200 cm3 |  |
| **H2SO4** | 4 | 2 dm3 |  |

**Molar volume of gases**

|  |
| --- |
| **1.35C understand how to carry out calculations involving gas volumes and the molar volume of**  **a gas (24 dm3 and 24,000 cm3 at room temperature and pressure (rtp)** |

In the case of a gas, there is another way, apart from molar mass of the gas, in which we can calculate how many moles of gas we have. We can do this because 1 mole of any gas, at room temperature and normal pressure (=rtp), occupies a volume of 24 dm3 which is called the molar volume of gases at room temperature.

So our next formula becomes:

|  |  |
| --- | --- |
| number of moles = | volume of gas in dm3 |
| 24 dm3 |

How many moles of gas are there in (at rtp):

1. 2400 cm3 of chlorine b. 120 cm3 of chlorine c. 96 cm3 of nitrogen d. 24 cm3 of methane

e. 12 dm3 of oxygen gas f. 12 dm3 of Cl2 g. 1.2 dm3 of SO2 h. 24 000 cm3 of O2

Calculate the volumes at RTP

a. 1.23 moles of N2 b. 2.05 moles of O2 c. 7 moles of NH3

d. 0.5 moles of CH4 e. 0.25 moles of He f. 0.001 moles of F2

**Section 1f: ionic compounds**

|  |
| --- |
| 1.37 understand how ions are formed by electron loss or gain  1.38 know the charges of these ions:   * metals in Groups 1, 2 and 3 * non-metals in Groups 5, 6 and 7 * Ag+, Cu2+, Fe2+, Fe3+, Pb2+, Zn2+ * hydrogen (H+), hydroxide (OH-), ammonium (NH4+), carbonate (CO32-), nitrate (NO3-), sulfate (SO42-) |

During many chemical reactions, atoms combine with each other by forming chemical bonds between them. During this bond making some or all of the reacting atoms re-arrange their valence electrons.

Atoms can bond in three different ways (or arrange their valence electrons in three different ways). The three different types of bonding between atoms are ionic bonds, covalent bonds and metallic bonds. Which type of bond is formed depends on the types of atoms that react. This is because the way in which atoms re-arrange their valence electrons depends on the electronic structure of the atoms involved in the reaction.

**Ionic bonds**

Ionic bonds are usually found in compounds that contain a metal (e.g. from groups 1, 2 or 3) that has combined with a non-metal (e.g. from groups 5, 6 or 7).

Ionic bonds are formed when an electron (or electrons) from the outermost shell of the metal atom is (are) transferred to the outermost shell of the non-metal atom.

In this process, the metal atom loses one or more electrons and the non-metal gains one or more electrons. An electron has been transferred from the metal to the non-metal.

Because of this transfer of electrons, each atom achieves the same electron configuration as the noble gas which is the nearest to it in the Periodic Table which is a full outer shell configuration or noble gas electronic configuration.

However, as a result of this transfer of electrons, the number of electrons and protons in the atoms is not equal anymore and the…

atoms become charged particles called ions because they have lost or gained electrons.

Ions are formed when atoms either gain or lose electrons and so become positively charged (by losing electrons) or negatively charged (by gaining electrons). During the formation of an ionic bond, atoms do not lose any protons or neutrons.

metals lose electrons, non-metals gain electrons!!

**Metals** Complete the table below showing how positive ions are formed

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **atom** | | | | | | **common ions** | | | |
| atom | number of protons (+) | number of electrons  (-) | group number | electronic structure | number of electrons transferred | number of protons (+) | number of electrons  (-) | electronic structure of ion | charge of ion |
| Na | 11 | 11 | 1 | 2,8,1 | loses 1 | 11 |  | 2,8 | Na+ |
| Mg |  |  |  |  |  |  |  |  |  |
| K |  |  |  |  |  |  |  |  |  |
| Ca |  |  |  |  |  |  |  |  |  |
| Al |  |  |  |  |  |  |  |  |  |

**Non-metals** Complete the table below showing how negative ions are formed.

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **atom** | | | | | | **common ions** | | | |
| atom | number of protons (+) | number of electrons  (-) | group number | electronic structure | number of electrons transferred | number of protons (+) | number of electrons  (-) | electronic structure of ion | charge of ion |
| F |  |  |  |  | gains 1 |  |  |  |  |
| Cl |  |  |  |  |  |  |  |  |  |
| O |  |  |  |  |  |  |  |  |  |
| S |  |  |  |  |  |  |  |  |  |

Study the table and answer the questions.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| substance | number of protons | number of neutrons | number of electrons |  | 1. Which is a positive ion? 2. Which is a negative ion? 3. Which is a metal ion? 4. Which is an isotope of C-12? |
| A | 9 | 10 | 10 |  |
| B | 12 | 12 | 12 |  |
| C | 19 | 20 | 18 |  |
| D | 6 | 8 | 6 |  |

Now that we know how ions are formed let’s now look at the actual transfers of electrons.

|  |
| --- |
| 1.40 draw dot and cross diagrams, the show the formation of ionic compounds by electron transfer,  limited to combinations of elements from Groups 1, 2, 3, and 5, 6, 7 |

Dot and cross diagrams allow us to determine the ratio’s in which positive and negative will bond; once we know that ratio we can write the formula of the ionic compound.

In the spaces below, using **x** and **0** to represent electrons from different atoms, show the formation of ionic bonds. Use square brackets for the ions. In each box also write the formula of the compound.

|  |
| --- |
| sodium fluoride (only draw outer electrons)  Na + F ⎯→ Na F    arrangement of electrons arrangement of electrons Formula of sodium chloride: in positive sodium ion in negative fluoride ion |
| potassium chloride  K + Cl ⎯→ K Cl  Formula: |

|  |
| --- |
| magnesium oxide  Mg + O ⎯→ Mg O  Formula: |
|  |
| calcium chloride  Ca + Cl ⎯→ Ca Cl  Formula: Cl |
|  |
| aluminium oxide  Formula: |
|  |
| magnesium nitride  Formula: |

**Naming ionic compounds**

|  |
| --- |
| * **if the compound contains TWO elements only**: |
| name of metal + first part of name of non-metal + **ide**  (or ammonium or hydrogen) |
| (e.g.: sodium bromide, hydrogen iodide, potassium fluoride, silicon carbide) |
|  |
| * **if compound contains THREE elements or more (one of the elements is usually oxygen):** |
| name of metal + first part of name of non-metal + **ate**  (or ammonium or hydrogen) |
| (e.g.: calcium nitrate, copper sulphate, sodium carbonate) |
|  |
| * the ion –OH- is called hydroxide whilst the ion NH4+ is called ammonium. Other common ions are carbonate = CO32- , nitrate = NO3-, sulphate = SO42-; phosphate = PO43-. |

Name the following compounds: a) KCl b) MgO c) CaCO3 d) LiOH e) NH4NO3

f) Ba(OH)2 g) Cs2 SO4 h) Fe2O3 I) Al2(SO4)3 j) Zn(NO3)2

**Writing formula of ionic compounds**

|  |
| --- |
| 1.39 write formulae for compounds formed between the ions listed above in statement 1.38 |

1. Write the symbol(s) for the atoms/ions/groups of atoms;

2. Write down their ionic charges

3. Balance the ionic charges (should add up to zero);

|  |  |  |
| --- | --- | --- |
| Examples: Write the formulae for ionic compounds | | |
| a. potassium fluoride | b. calcium iodide | c. barium nitrate |
| d. aluminium oxide | e. ammonium sulphate | f. iron (II) hydroxide |

|  |  |  |
| --- | --- | --- |
| Exercise: Write the formulae of the following ionic compounds | | |
| a. sodium fluoride | b. magnesium iodide | c. ammonium hydroxide |
| d. beryllium chloride | e. sodium oxide | f. iron (II) nitrate |
| g. aluminium chloride | h. copper (II) carbonate | i. ammonium phosphate |
| j. caesium nitrate | k. potassium phosphate | l. silver (I) nitrate |
| m. lead (II) chloride | n. barium sulphate | o. calcium hydroxide |
| p. calcium phosphate |  |  |

**Table of valencies of common ions**

valency = number of electrons given away/accepted/shared

|  |  |  |  |
| --- | --- | --- | --- |
| positive | | negative | |
| species | valency | species | valency |
| Li | +1 | Cl | -1 |
| Na | +1 | Br | -1 |
| K | +1 | I | -1 |
| Mg | +2 | O | -2 |
| Ca | +2 | N | -3 |
| Al | +3 | SO42- | -2 |
| Fe | +2/+3 | CO32- | -2 |
| Cu | +2 | NO3- | -1 |
| NH4+ | +1 | OH- | -1 |

## *(carbon, silicon = 4) (PO43- = -3)*

**Physical properties of ionic compounds**

|  |
| --- |
| 1.41 understand ionic bonding in terms of electrostatic attractions  1.42 understand why compounds with giant ionic lattices have high melting and boiling points  **1.43 know that ionic compounds do not conduct electricity when solid, but do conduct**  **electricity when molten and in aqueous solution.** |

**An ionic bond = the electrostatic attraction between a positive ion (usually a metal ion) and a negative ion (a non metal ion).**

As these electrostatic attractions between the oppositely charged are usually strong, ionic compounds have high melting and boiling points. This is because the ions need a lot of energy to overcome the strong electrostatic attraction between them so they can move away from their positions in giant ionic lattice or to escape from the liquid when boiling.

Giant ionic lattice

The ions in an ionic compound such as sodium chloride are arranged in a giant three dimensional **lattice** that is called a **giant ionic lattice**; a lattice has a pattern that is repeated in all directions.

|  |  |
| --- | --- |
|  | This regular arrangement results in the formation of a crystal that is held together by the **strong electrostatic attraction between the oppositely charged ions.** Apart from being strong forces, they act in all directions.  The lattice on the left could represent the lattice of sodium chloride in a sodium chloride crystal. The positive particles are sodium ions whilst the negative ones are **chloride ions.** |

|  |  |
| --- | --- |
| Because of the giant ionic lattice ionic compounds have the following properties: | * high melting and boiling points because of the strong attraction between the ions that act in all directions * soluble in water * conduct electricity when molten and dissolved in water but not as solids |

For a substance to be able to conduct electricity it needs to have charged particles that are mobile. In the case of ionic solids the charged particles are present, i.e. ions, but they are held in fixed positions in the giant lattice and are unable to move; as a result ionic solids cannot conduct electricity.

Heating or dissolving ionic solids allows the ions to move freely so that electricity is conducted.

**Section 1g: covalent compounds**

|  |
| --- |
| 1.44 know that a covalent bond is formed between atoms by sharing of a pair of electrons  1.45 understand covalent bonds in terms of electrostatic forces between the bonding pair of electrons  and the nuclei of the atoms involved in the bond |

**Covalent bond**

Another way for atoms to obtain noble gas electronic configuration or a full outer shell is by sharing valence electrons with another atom or atoms. This sharing of electrons happens when non-metal atoms of the same or of different elements react with each other.

Atoms can share valence electrons (=electrons in their outer shell) by allowing their outer shells to overlap. In a single covalent bond each atom gives 1 electron to form a pair of shared electrons. The shared electrons become part of the outer shell of both atoms involved in the bond. The pair of electrons is used to bond both atoms i.e. to ensure that they are attracted to each other.

When two atoms overlap so they can share a pair of electrons a **covalent bond** is formed. **A covalent bond is a shared PAIR of electrons**. The pair of electrons which are shared by both atoms are also called a bonding pair as they form a covalent bond. A covalent bond is a strong attraction between the shared or bonding pair of electrons and the nuclei of **both** atoms that share their electrons.

The particle formed by this type of bonding is a **molecule**. Covalent compounds are compounds which are formed when atoms share pairs of electrons.

Some atoms need more than one electron to fill their outer shell. In that case they will form a double bond (if 2 electrons are needed) or even a triple bond (if 3 more electrons are needed). In the case of a double bond, an atom will share two of its valence electrons and in the case of a triple bond, the atoms will share 3 electrons.

|  |
| --- |
| 1.46 understand how to use dot and cross diagrams to represent covalent bonds in:   * diatomic molecules, including hydrogen, oxygen, nitrogen, halogens and hydrogen halides * inorganic molecules including water, ammonia and carbon dioxide * organic molecules containing up to two carbon atoms, including methane, ethane, ethane and those containing halogen atoms |

For each of the above molecules draw a diagram to show the arrangement of the valence electrons i.e. show how the electrons are shared between the atoms in the molecule.

A few simple rules:

* the number of valence electrons usually tells us how many covalent bonds an element will form:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Element | Group number | Number of valence electrons | How many electrons short of octet | Number of covalent bonds the atom usually makes in a covalent compound | How many valence electrons are used in bonding |
| carbon | 4 | 4 | 4 | 4 | 4 |
| nitrogen | 5 | 5 | 3 | 3 | 3 |
| oxygen | 6 | 6 | 2 | 2 | 2 |
| sulphur | 6 | 6 | 2 | 2 | 2 |
| chlorine | 7 | 7 | 1 | 1 | 1 |
| iodine | 7 | 7 | 1 | 1 | 1 |
| hydrogen |  | 1 | 1 | 1 | 1 |

* if there is more than 2 atoms in the molecule, the atom which makes the highest number of covalent bonds usually goes in the middle of the molecule.

There are two types of covalent substances i.e. substances with covalent bonds:

* covalent substances with a simple molecular structure
* covalent substances with giant molecular structures

The structure determines the physical properties of the compound.

**Simple molecular structures**

|  |
| --- |
| 1.47 explain why substances with simple molecular structures are gases or liquids, or solids with low  melting points  1.48 explain why the melting and boiling points of substances with simple molecular structures  increase, in general, with increasing relative molecular mass. |

**Simple molecular structure** refers to a structure in which small molecules are attracted to each other by weak intermolecular forces. Within the molecules the atoms are held together by covalent bonds which are strong.

Examples of substances with simple molecular structures are all the substances for which you drew cross and dot diagrams on page 1. Others substances are iodine, carbon monoxide, glucose, ethanol, …

The physical properties of simple molecular substances are influenced by the **weak** forces of **attraction between** the molecules. As a result **molecules need little energy to overcome** the forces of attraction and increase their motion.

|  |  |
| --- | --- |
| Simple molecular substances have the following properties:   1. low melting and boiling points which is why many molecular compounds are gases or liquids 2. those which are solids are volatile i.e. they change state easily because they have low melting and boiling points;   As molecular substances do not have any charged particles such as electrons or ions that can move through the substance, they do not conduct electricity. |  |

**Giant covalent structures**

|  |
| --- |
| 1.49 explain why substances with giant covalent structures are solids with high melting and boiling points.  1.50 explain how the structures of diamond, graphite and C60 fullerene influence their physical properties,  including electrical conductivity and hardness  1.51 know that covalent substances do not usually conduct electricity |

A giant covalent structure refers to a structure in which a very large number of atoms are held together by a very large number of strong covalent bonds that extend in all directions. There are no weak intermolecular forces in this structure.

The melting and boiling of such substances requires a lot of energy for the **atoms** to break the covalent bonds which are strong and of which there are many. As a result giant covalent structures are solids with very high melting and boiling points.

*Examples of giant covalent structures*

Diamond, graphite and C60 fullerenes are different forms in which carbon can exist in nature. Diamond, graphite and fullerenes are referred to as allotropes; allotropes are different forms of the same element.

The structure of diamond and graphite has the following characteristics:

* the lattice has carbon atoms only – there are no molecules in it;
* the atoms are held together by strong covalent bonds
* the lattice has millions of carbon atoms going in all directions.

Both diamond and graphite have a giant covalent structure = a structure in which a very large number of separate atoms are held together by strong covalent bonds only. **Giant structure means that the bonding goes on and on to the edge of the crystal.**

It is different from a simple molecular structure, in which, although the bonds between the atoms in a molecule are strong covalent bonds, the intermolecular forces of attraction are weak.

The strong covalent bonds explain the hardness and high melting and boiling points of giant covalent substances!

*How are diamond and graphite different?*

Both are giant covalent solids but differ in some properties because of an important structural difference.

|  |  |
| --- | --- |
| diamond | graphite |
|  |  |
| The carbon atoms in diamond form a three-dimensional structure: each carbon atom in the diamond lattice has made strong covalent bonds with 4 other carbon atoms | In the graphite lattice each carbon has made covalent bonds with only 3 atoms which are all in the same layer. Graphite has a layer structure with the carbon atoms in the same layer bonded to each other by strong covalent bonds. The layers themselves are held together by weak intermolecular forces. |

Because of these differences in structure, graphite and diamond have the following different properties and uses:

|  |  |
| --- | --- |
| Difference | Reason |
| Diamonds are hard which is why diamonds are used for cutting  Graphite is soft which is why it used as a lubricant (=substance used to reduce friction) | * Diamond is used for cutting because all atoms are held together by strong covalent forces (there is no other force in the lattice); covalent bonds between all atoms * Graphite is used as a lubricant because it is soft and slippery because the carbon atoms in graphite are in layers that easily slip past each other as they are held in place by weak forces of attraction. |
| **Covalent substances do usually not conduct electricity**  Usually covalent substances do not conduct as they do not have charged particles such as electrons and ions that can move freely which is why diamond is an insulator; however, not all covalent substances are insulators. | |
| Diamond is an insulator whilst  Graphite conducts | * There are no mobile charged particles * each carbon atoms has only made 3 covalent bonds and has therefore only used 3 of its 4 valence electrons; the valence electrons that has not been used to make covalent bonds can move freely throughout the lattice (delocalised electron) |

**C60 fullerene – not really a giant covalent structure**

|  |  |
| --- | --- |
| There is another allotrope of carbon that is called C60 fullerene.  It consists of molecules (C60) made up of 60 atoms bonded covalently and arranged in the form of a football. C60 fullerene have a **simple molecular structure**. Between the molecules there are weak forces of attraction which is the reason fullerenes sublime easily.  In C60 fullerene molecules each carbon has, just like in graphite, only made 3 covalent bonds and has therefore also delocalized electrons and can conduct electricity. However, it is not so easy for the electrons to move between the molecules (spheres) and it is therefore not as good a conductor as graphite |  |

Other examples of giant covalent structures are silicon and silicon dioxide.

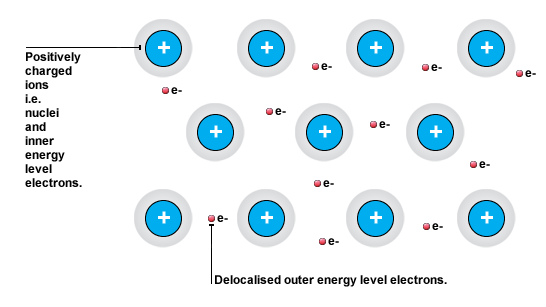
|  |  |
| --- | --- |
| silicon | silicon dioxide – SiO2 |
| The spheres above all represent silicon atoms bonded covalently. | The red spheres represent oxygen as it is they have made 2 covalent bonds – oxygen only needs to share 2 electrons. The black spheres are silicon as they have 4 bonds – silicon needs to share 4 electrons. |

**Section 1h: metallic bonding**

|  |
| --- |
| 1.52 know how to represent a metallic lattice by a 2-D diagram  1.53 understand metallic bonding in terms of electrostatic attractions  1.54 explain typical physical properties of metals, including electrical conductivity and malleability |

Metallic crystals have a giant structure of positive ions surrounded by a ‘sea’ of delocalised electrons which can move around in the lattice as shown below. The bond between the metal ions is called a metallic bond. A metallic bond is the attraction between the positive metal ions and the delocalized electrons.

(from <http://www.bbc.co.uk/scotland/learning/bitesize/higher/chemistry/energy/bsp_rev3.shtml> on 15/6/2010)



This structure and bonding allows the following properties as shown by the table below.

|  |  |
| --- | --- |
| **Property** | **Explanation** |
| Metals are **ductile** (=can be pulled into wires) and **malleable** (= can be bent or hammered into different shapes without breaking);  . | The layers of positive metal ions can slide past each other without breaking the metallic bond; this is the case because there will always be delocalised electrons between these positive metal ions even when they change their positions. There will always be a force of attraction i.e. the metallic bonds to hold the metal ions together. |
| Metals conduct electricity and heat well. | Sea of delocalised electrons can move through the lattice and carry the negative charge through lattice. |

**Section 1i: electrolysis**

**Electrolytes**

|  |
| --- |
| **1.55C understand why covalent compounds do not conduct electricity**  **1.56C understand why ionic compounds conduct electricity only when molten or in aqueous**  **solution**  **1.57C know that anion and cation are terms used for negative and positive ions respectively** |

In this topic we will study the effect of passing an electric current through a compound.

An electric current is a flow of electrons. For a substance to conduct electricity it must have mobile charged particles i.e. electrons or ions. For instance metals can conduct electricity (a current passes through them) because metals have delocalized electrons that can move through the metal.

Covalent compounds do not have any charged particles as they consist of atoms or molecules that are electrically neutral; as a result they cannot conduct electricity even if we melt or dissolve them in water (if they are soluble that is).

On the other hand ionic compounds have charged particles (metal positive ions and negative non-metal ions). However, ionic compounds can only conduct when the ionic compound is molten or dissolved in water. Only in the liquid or aqueous state can the ions (charged particles) move to the oppositely charged electrode.

A liquid (molten or solution) that conducts electricity is called an **electrolyte.**

We can tests substances if they are electrolytes or non-electrolytes (liquids which do not conduct electricity) by doing simple tests.

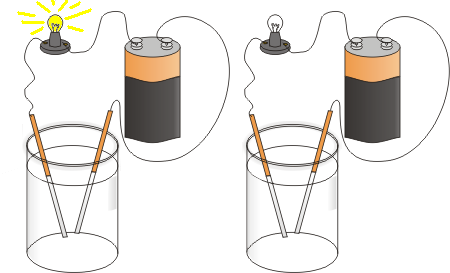
A circuit is set up as shown below. It consists of a battery or power supply, wires, carbon electrodes and an indicator of a current i.e. a bulb or an ammeter. Carbon is picked as a material for the electrodes because it is **inert** i.e. it does not react and does therefore not interfere in the electrolysis reaction. The electrode which is connected to the positive terminal of the battery is named the **anode**; the electrode connected to the negative one is the **cathode.**

The substance to be tested is either molten or dissolved in water. If the bulb lights or a current is measured by an ammeter, the substance tested is an electrolyte. In the diagram below the compound on the left is ionic – the bulb is on - whilst the one on the right is covalent as the bulb does not light.

covalent compound

molten or dissolved ionic compound

(electrolyte)



(image from <http://www.cdli.ca/courses/sci1206/unit03_org02_ilo06/b_activity.html> on 22/06/2010)

**Electrolysis**

|  |
| --- |
| **1.58C describe experiments to investigate electrolysis, using inert electrodes, of molten**  **compounds (including lead (II) bromide and aqueous solutions (including sodium chloride,**  **copper(II) sulphate and dilute sulphuric acid) and predict the products**  **1.59C write ionic half-equations representing the reactions at the electrodes during electrolysis**  **and understand why these reactions are classified as reduction and oxidation** |

An electrolyte is a different type of conductor than a solid metal. When it conducts, the current passing through the electrolyte causes a chemical reaction to take place i.e. new substances are formed. The ionic compound which makes up the electrolyte is decomposed. A reaction in which a compound is decomposed using an electrical current is called an **electrolysis** reaction; as a result new substances are formed when an ionic compound conducts electricity. The current causes a chemical reaction converting electrical energy into chemical energy.

You need to be able to predict the products of electrolysis reactions. To be able to do this you need to know:

* what the ionic compound is
* if it is molten or dissolved in water.

**Electrolysis of molten ionic compounds or salts**

When a molten salt is electrolyzed, the positive metal ions or **cations** are attracted to the negative electrode or cathode where they lose their charge as they gain electrons (**reduction**) and become metal atoms. The negative non-metal ions or **anions** are attracted to the anode where they lose their negative charge. They lose electrons (**oxidation**) and become non-metal atoms again and usually form molecules.

This is shown by the example below which shows the electrolysis of molten lead bromide. The products are bromine and lead.

The actual changes going on at each electrode can be shown in ionic half equations:

|  |  |  |
| --- | --- | --- |
| Half equation at the cathode (-)  Pb2+(l) + 2e- → Pb (s)  reduction |  | Half equation at the anode (+)  2Br-(l) → Br2 (l) + 2e-  oxidation |
|  | | |

The flow of charge is maintained in the following way:

* by the flow of electrons in the wires and the electrodes and
* by the movement of ions towards the oppositely charged electrode in the electrolyte.

***When a molten salt is electrolyzed the metal is produced at the cathode whilst the non-metal is formed at the anode.***

**Electrolysis of sodium chloride solution**

|  |  |
| --- | --- |
|  | When a solution is electrolyzed predicting the products becomes more complicated as also the water (the solvent) also becomes involved. Water, although a covalent compound, ionizes a little bit producing hydrogen and hydroxide ions which are also attracted to the electrodes. In most cases a gas will be produced which needs to be collected and tested. When carrying out electrolysis during which a gas will be produced the following set up ca be used to collect the gas(es). |

|  |  |
| --- | --- |
| Starting chemicals | * concentrated sodium chloride solution or brine, NaCl (aq) * ions present: * at cathode: Na+ H+ at the anode: Cl- OH- * only 1 ion can react or be discharged at each electrode, the other remains in solution |
| observations | **anode**   * bubbles, yellow green gas, nearly 1:1 ratio with gas at cathode * test: gas bleaches damp litmus paper |
| **cathode**   * bubbles, colourless gas * test: ‘pop’ sound with lit splint, more than 1:1 ratio * red litmus turns blue |
| products | **anode**   * chlorine is discharged and not the hydroxide ions which remain in solution * ionic half equation: 2Cl- (aq) → Cl2 (g) + 2e-  = oxidation |
| **cathode**   * hydrogen is discharged:   + ionic half equation: 2H+(aq) + 2e- → H2 (g) = reduction * sodium hydroxide: Na+ and OH- are left in the solution as sodium hydroxide which can be detected using an indicator |
| overall equation | at anode: 2Cl- (aq) ⎯→ Cl2 (g) + 2e-  at cathode: 2H+ (aq) + 2e- ⎯→ H2(g)  2H+(aq) +2Cl- (aq) ⎯→ Cl2 (g)+ H2 (g) |

Trend: for any concentrated solution of a salt made from a group 1 metal and a group 7 non-metal there will always be three products: (1) hydrogen, (2) the halogen, and (3) the hydroxide of the group 1 metal.

**Electrolysis of copper (II) sulphate solution using inert electrodes**

|  |  |
| --- | --- |
| starting chemicals | * copper sulphate solution, CuSO4 (aq) * ions present: at cathode: Cu2+ H+ at the anode: SO42- OH- |
| observations | **anode**   * bubbles, colourless gas * test: relights glowing splint |
| **cathode**  red deposit on electrode |
| products | **anode**   * oxygen is produced * ionic half equation: 4OH- (aq) → O2 (g) + 2H2O (l) + 4e- = oxidation * sulphate, SO42- remains in solution |
| **cathode**   * copper   + is discharged and not the hydrogen ions as they remain in the solution:   + ionic half equation: Cu2+ (aq) + 2e- → Cu (s) = reduction * hydrogen ions remain in solution |
| overall equation | at anode: 4OH- (aq) → O2 (g) + 2H2O (l) + 4e-  at cathode: 2Cu2+ (aq) + 4e- → 2Cu (s)  4OH- (aq) + 2Cu2+ (aq) → O2 (g) + 2H2O (l)+ 2Cu (s) |

**Electrolysis of dilute sulphuric acid**

|  |  |
| --- | --- |
| The two previous solutions are considered concentrated solutions. The electrolyte, sulphuric acid, in this example is diluted and this also has an effect on kind of products of the electrolysis as none of the ions of the electrolyte will be discharged. Instead they remain in the solution. The ions which are discharged are the hydrogen and hydroxide ions which form hydrogen gas and oxygen gas.  In essence when a dilute solution is electrolyzed water is decomposed. The identity and nature of the electrolyte does not really matter | Diagram of the Hofmann Voltameter to collect gases from electrolysis of acid or salt solutions (c) doc b |

**(above image from** [**http://www.docbrown.info/page01/ExIndChem/ExtraElectrochem.htm**](http://www.docbrown.info/page01/ExIndChem/ExtraElectrochem.htm) **on 21/06/2010)**

|  |  |
| --- | --- |
| starting chemicals | * dilute sulphuric acid, H2SO4 (aq) * ions present: at cathode: H+ at the anode: SO42- OH- |
| observations | **anode**   * bubbles, colourless gas, * volume approximately half of the gas produced at the cathode * test: relights glowing splint |
| **cathode**   * bubbles, colourless gas * volume: twice as much as gas at the anode * test: ‘pop’ sound with lit splint, |
| products | **anode**   * oxygen is produced * ionic half equation: 4OH- (aq) → O2 (g) + 2H2O (l) + 4e-  = oxidation |
| **cathode**   * hydrogen is discharged: * ionic half equation: 2H+(aq) + 2e- → H2 (g) = reduction |
| overall equation | at anode: 4OH- (aq) → O2 (g) + 2H2O (l) + 4e-  at cathode: 4H+ (aq) + 4e- → 2H2 (g)  4OH- (aq) + 4H+ (aq) → O2 (g) + 2H2O (l)+ 2H2 (g) |

**Electrolysis of copper solution using copper electrodes as used in the purification of copper.**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | |  |  | | --- | --- | | cathode:  Cu2+ (aq) + 2e- ⎯→ Cu( s) | anode:  Cu (s) ⎯→ Cu2+ (aq) + 2e- | | Observation  Copper ions from solution are  deposited onto piece of pure copper.  Mass of electrode increases | Observation  Copper from impure copper goes  into solution; impurities sink.  Mass of electrode decreases | |

**Section 2a: The Group 1 elements — lithium, sodium and potassium**

|  |
| --- |
| 2.1 understand how the similarities in the reactions of these elements with water provide evidence for  their recognition as a family of elements  2.2 understand how the differences between the reactions of these elements with air and water provide  evidence for the trend in reactivity in Group 1  2.3 use knowledge of trends in Group 1 to predict the properties of other alkali metals  **2.4C explain the trend in reactivity in Group 1 in terms of electronic configurations** |

Similarities: Lithium, sodium and potassium have the following similar properties:

1. they are soft as you can cut them easily with a knife;
2. low densities (why they float);
3. conduct electricity and heat as they are metals – metallic bonding.
4. reactions:
5. react with water to form a metal hydroxide which dissolves (= alkali) and hydrogen;
6. react with air to form a basic metal oxide that dissolves to form a metal hydroxide solution (=alkali);
7. stored under paraffin as they react readily with oxygen and water from the atmosphere

It is because of the similarities in their chemical behaviour that they are considered a family of elements.

You need to know these equations well

2Li (s) + 2H2O (g) ⎯→ 2LiOH (aq) + H2 (g)

2Na (s) + 2H2O (g) ⎯→ 2NaOH (aq) + H2 (g)

2K (s) + 2H2O (g) ⎯→ 2KOH (aq) + H2 (g)

|  |  |
| --- | --- |
| Observations when Na, Li and K react with water:   * they float * move across the surface of the water, * fizz - produce gas | * potassium catches fire (lilac flame) * the solution has blue colour when universal indicator is added = alkali * potassium and sodium melt to form a sphere |

**Differences (trends) within the group**

A trend is a gradual change. We know which properties the group 1 metals share however….

… as you go down the group 1:

* reactivity with water and air (but also with other chemicals) increases: potassium (of Li, Na and K) is the most reactive and lithium is the least reactive. In the whole group, francium is the most reactive.
* melting and boiling point decrease: lithium has highest melting and boiling point.
* densities increase.

**Predicting properties of the other group 1 elements**

Using the trends described above we should be able predict the properties of the remaining group 1 elements rubidium, caesium and francium.

Complete the table below

|  |  |  |  |
| --- | --- | --- | --- |
| Group 1 element | Melting point (0C) | Density (g/cm3) | Reaction with water |
| Li | 181 | 0.535 | Fizzes steadily, gradually disappears |
| Na | 98 | 0.971 | Fizzes rapidly, melts into a ball and disappears quickly |
| K | 63 | 0.862 | Ignites easily, sparks, lilac flame and disappears very quickly |
| Rb |  | 1.53 | Explodes with sparks |
| Cs | 29 |  |  |

**Why does reactivity increase as you go down group 1?**

When group 1 metals react they lose the single valence electron – they oxidize. However, the negative valence electron is attracted by the positive nucleus.

As you go down the group the atoms of the group 1 elements become larger i.e. potassium has the largest atoms of the three (Li, Na and K); this also means that the valence electron in a potassium atoms is the furthest away from the nucleus and is attracted the least strongly. This is why potassium more easily loses its valence electron; this is why potassium is more reactive than sodium and lithium and sodium is more reactive than lithium.

**Section 2b: The Group 7 elements — chlorine, bromine and iodine**

|  |
| --- |
| 2.5 know the colours, physical states (at room temperature) and trends in physical properties of these  elements  2.6 use knowledge of trends in Group 7 to predict the properties of other halogens  2.7 understand how displacement reactions involving halogens and halides provide evidence for the  trend in reactivity in Group 7  **2.8C explain the trend in reactivity in Group 7 in terms of electronic configurations** |

**Appearance and physical states of the halogens**

Study the trends in properties and complete the table.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **element** | **colour** | **state** | **density (g/cm3)** | **melting point (in C)** |
| fluorine |  |  | 0.001696 | -219 |
| chlorine | yellow green | gas | 0.003214 | - 101 |
| bromine | red brown | liquid | 3.122 | -7.2 |
| iodine | dark grey | solid | 4.809 | 113.5 |
| astatine |  |  | 7 | 302 |

**Another family of elements**

Group 7 elements share the following properties:

1. they are all diatomic molecules i.e. F2, Cl2, Br2 and I2 , At2.
2. react with:
3. metals to form ionic compounds called salts
4. water to form acidic solutions
5. hydrogen to form hydrogen halides such as HF, HCl, HBr and HI.

**Trends within the group**

… as you go down group 7 …..

reactivity decreases; any halogen displaces a halogen ion from its salt which is lower in the group as shown by displacement reactions.

**Displacement reactions between halogens and halides (compounds with halogens)**

These are reactions between solutions of the halogens (Cl2, Br2 and I2 are only slightly soluble in water but soluble enough to make a solution) and solutions of their ionic compounds or salts or halides. To be able to analyze experimental observations the following information is needed:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| appearance of halogens dissolved in water | | appearance of halogens dissolved in hexane | | appearance of sodium halides | |
| Cl2 (aq) | colourless | Cl2 (aq) | colourless | NaCl (aq) | colourless |
| Br2 (aq) | orange | Br2 (aq) | red brown | NaBr (aq) |
| I2 (aq) | brown | I2 (aq) | purple | NaI (aq) |

Halogens prefer to be dissolved in hexane instead of water.

The table below summarises the results of a set of displacement experiments.

|  |  |  |  |
| --- | --- | --- | --- |
| **halogen** | **solutions of halogen salts (halides) to which halogen is added** | | |
| NaCl (aq)  (colourless) | NaBr (aq)  (colourless) | NaI (aq)  (colourless) |
| Cl2 (aq)  (colourless) | no colour change | colourless solution goes orange (bromine is produced) | colourless solution goes red-brown  (iodine is produced) |
| Br2 (aq)  (orange) | no colour change | no colour change | colourless solution goes red-brown  (iodine is produced) |
| I2 (aq)  (red-brown) | no colour change | no colour change | no colour change |

**Redox reactions.** The above observations indicate that the following reactions have taken place:

|  |
| --- |
| Cl2 (aq) + 2NaBr (aq) → Br2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces the bromine. The chlorine atoms are reduced (each gains an electron) and bromide ions are oxidized (each lose an electron). This displacement reaction is also a redox reaction:  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2Br- → Br2 +2e-  Ionic equation: Cl2 (aq) + 2Br- (aq) → Br2 (aq) + 2Cl- (aq) |
| Cl2 (aq) + 2NaI (aq) → I2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each loses an electron).  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2I- → I2 +2e-  Ionic equation: Cl2 (aq) + 2I- (aq) → I2 (aq) + 2Cl- (aq) |
| Br2 (aq) + 2NaI (aq) → I2 (aq) + 2NaBr (aq)  During this reaction, the bromine displaces the iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each lose an electron).  Half equations: reduction: Br2 + 2e- → 2Br- oxidation: 2I- → I2 +2e-  Ionic equation: Br2 (aq) + 2I- (aq) → I2 (aq) + 2Br- (aq) |

**Section 2c: Gases in the atmosphere**

**Oxygen in air**

|  |
| --- |
| 2.9 know the approximate percentages by volume of the four most abundant gases in dry air  2.10 understand how to determine the percentage by volume of oxygen in air using experiments  involving the reactions of metals (e.g. iron) and non-metals (e.g. phosphorus) with air  2.14 practical: determine the approximate percentage by volume of oxygen in air using a metal or a  non-metal |

The table shows the composition in percentage in volume of the main gases in a sample of clean air.

|  |  |
| --- | --- |
| **gas** | **composition by volume (%)** |
| nitrogen | 78 |
| oxygen | 21 |
| carbon dioxide | 0.04 |
| argon | 0.9 |

The percentage by volume of oxygen in air can be determined by the following experiments:

Experiment with copper

Watch the animation on <http://www.bbc.co.uk/schools/gcsebitesize/science/aqa/oils/changesrev5.shtml>

(the link is also on the wiki) and answer the following questions.

The experiment involves the following setup:



1. Complete the results table.

|  |  |
| --- | --- |
| Volume of air at the start (cm3) | Volume at the end when reaction is complete (cm3) |
|  |  |

1. What caused the decrease in the volume?
2. Write a word equation and balanced symbol equation of the reaction that is taking place.
3. Describe the appearance of the product.
4. Why is the apparatus allowed to cool before the final volume of air is measured?

6. Identify any experimental errors.

Experiment with iron

|  |  |
| --- | --- |
| A piece of iron wool is placed in a test tube which is inverted in a beaker of water. The test tube is filled with air and is left for a few days to allow the iron to rust. The length of the column of air in the test tube is measured at the start of the experiment and again when no more change to the length of the column or the iron wool is observed. The diagram below shows the test tube at the start and at the end of the experiment. |  |

|  |  |
| --- | --- |
| 1. Describe the observations at the end of the experiment. Make a diagram in the box to the right to show the length of the column of air at the end of the experiment. 2. Explain why the length of the column decreased by the amount observed. 3. How does the experiment provide evidence that air is 21% oxygen by volume? 4. Identify any experimental errors. |  |

Experiment with phosphorus

A marked bell jar with its lid removed is placed on top of coins inside a trough. The coins ensure that water can go inside the bell jar. Water is poured into the through until it reaches level zero marked on the bell jar. A piece of phosphorus is placed in an evaporating dish which floats on the water.

The phosphorus is lit and a lid placed at the top of the bell jar to cover it.

The yellow phosphorus starts burning, and the air space inside gets filled with white fumes. The white fumes are phosphorus oxide that dissolves in the water forming an acidic solution. When the phosphorus has stopped burning and the white fumes have disappeared the level of water in he bell jar is measured.

|  |  |
| --- | --- |
|  |  |
|  |  |

(from <http://www.tutorvista.com/content/chemistry/chemistry-i/air-mixture/air-oxygen.php> on 21/1/11)

1. Why did the water level go up in the bell jar?
2. What evidence does this experiment provide?
3. Identify any experimental errors?

**Combustion products: oxides**

|  |
| --- |
| 2.11 describe the combustion of elements in oxygen, including magnesium, hydrogen and sulfur |

Oxygen produces the following oxides during combustion reactions. Complete the last column.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Element** | **Description** | **Equation** | **pH of solution of oxide in water** | **Acid-base character of oxide** |
| Mg | White bright flame, forming a white solid called magnesium oxide | 2Mg (s) + O2 (g) → 2MgO (s) | 8 |  |
| C | Colourless and odourless gas produced | C (s) + O2 (g) → CO2 (g) | 6 |  |
| S | Burns with a blue flame; pungent, poisonous yellowish gas produced | S (s) + O2 (g) → SO2 (g) | 1 |  |
| H | Exothermic reaction (could be explosive) producing steam; hydrogen burns with a light blue flame. | 2H2 (g) + O2 (g) → 2H2O (l) | 7 |  |

**Carbon dioxide**

|  |
| --- |
| 2.12 describe the formation of carbon dioxide from the thermal decomposition of metal carbonates,  including copper(II) carbonate  2.13 know that carbon dioxide is a greenhouse gas and that increasing amounts in the atmosphere  may contribute to climate change |

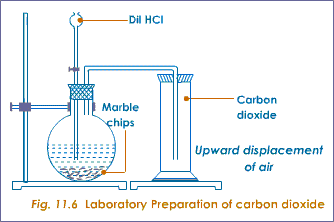
There are two ways in which carbon dioxide can be prepared in the laboratory:

Metal carbonate + acid (see 2f)

Carbon dioxide can be prepared using calcium carbonate or marble and dilute hydrochloric acid as shown by the equation below.

CaCO3 (s) + 2HCl (aq) → CaCl2 (aq) + 2H2O (l) + CO2 (g)

Carbon dioxide can be collected over water as it has a low enough solubility.



(from <http://www.tutorvista.com/content/chemistry/chemistry-i/air-mixture/air-oxygen.php> on 21/1/11)

Thermal decomposition of a metal carbonate

When copper (II) carbonate is heated it quickly decomposes (copper is a less reactive metal) into copper oxide and carbon dioxide. The apparatus is shown below.

|  |  |
| --- | --- |
|  | 1. Write a balanced symbol equation for this reaction. 2. Write down observations you would expect to make during this experiment. |

(from <http://www.practicalchemistry.org/experiments/thermal-decomposition-of-metal-carbonates,281,EX.html> on 21/1/11)

Greenhouse gas and climate change

The Earth absorbs a lot of the energy it receives from the sun but it also reradiates some of that energy as infrared radiation back into the atmosphere and space. Carbon dioxide in the atmosphere absorbs some of the infrared radiation. As a result not all of the heat energy goes back to space; this keeps the Earth warm. Gases which absorb the infrared radiation that the Earth reradiates are called greenhouse gases.

Because of deforestation and increased burning of fossil fuels there is now more carbon dioxide in the atmosphere than ever before. Evidence suggests (not conclusively) that as a result of this, the Earth is gradually warming (on average as not necessarily everywhere) and this causes climate change as some places on Earth are getting hotter, others are getting colder, or drier or windier or experience more destructive hurricanes or more severe drought.

**Section 2d reactivity series**

|  |
| --- |
| 2.15 understand how metals can be arranged in a reactivity series based on their reactions with:  • water • dilute hydrochloric or sulfuric acid.  2.16 understand how metals can be arranged in a reactivity series based on their displacement  reactions between: metals and metal oxides and metals and aqueous solutions of metal salts.  2.17 know the order of reactivity of these metals: potassium, sodium, lithium, calcium, magnesium,  aluminium, zinc, iron, copper, silver, gold.  2.21 *practical: investigate reactions between dilute hydrochloric and sulfuric acids and metals (e.g.*  *magnesium, zinc and iron)* |

**Reaction of metals with cold water**

|  |  |
| --- | --- |
| ***Observations:*** the metals that react with cold water are: all group 1 metals, calcium and magnesium (very slowly). Potassium catches fire. | ***Conclusion:*** only reactive metals react with water and this reaction produces an hydroxide and hydrogen as shown by the symbol equations below - the rate at which hydrogen is released can be used to compare reactivity. |

Write balanced symbol equations for the reactions of (i) sodium and water and (ii) lithium with water, include state symbols.

|  |
| --- |
| (i) |
| (ii) |

**Reaction of metals with steam**

|  |  |
| --- | --- |
| Those metals that do not react with cold water can be tested with steam.  Reactive metals react very vigorously with steam therefore steam should only really be used with those metals that show a low rate of reaction with cold water or no reaction at all.  A possible set up for a reaction with steam is shown in the diagram to the right. |  |

(from <http://cbse.myindialist.com/chemistry-x-metals-and-non-metals-what-happens-when-metals-react-with-water/> on 21/1/11)

**Reaction of metals with dilute hydrochloric acid**

Most metals also react with acids. When they do they produce a salt and hydrogen according to the general equation:

metal + hydrochloric acid ⎯→ metal chloride + hydrogen

Write a word equation and a balanced symbol equation for the reaction of calcium with hydrochloric acid and zinc with hydrochloric acid. Include state symbols.

1. calcium + hydrochloric acid:

2. zinc + sulphuric acid:

**Summary of the reactions of metals with water/steam and acid.**

|  |  |  |  |
| --- | --- | --- | --- |
| **metal** | **reaction with cold water** | **reaction with steam** | **reaction with hydrochloric acid** |
| K | fizzes, violently, moves around, catches fire immediately, lilac flame | reaction with steam is too violent | very dangerously explosive |
| Na | fizzes, violently, moves around, might catch fire | dangerously explosive |
| Li | fizzes, moves around | very fast reaction, can be explosive |
| Ca | fizzes slowly | fizzes fast, hydrogen given off |
| Mg | reacts very slowly | reacts to form hydrogen and magnesium oxide | fizz, hydrogen given off |
| Al | no immediate reaction with cold water | reacts to form hydrogen and aluminium oxide | fizz, hydrogen given off |
| Zn | no immediate reaction with cold water | reacts to form hydrogen and an zinc oxide | fizz, hydrogen given off |
| Fe | no immediate reaction with cold water | reacts to form hydrogen and iron oxide | fizz, hydrogen given off |
| Cu | no reaction with cold water | no reaction | no reaction as it cannot displace hydrogen from an acid |

According to the observations in the table above, potassium is the most reactive metal of the metals included in the table and copper is the least reactive. We also know of course that gold and silver are even less reactive than copper.

To summarize further: the order in which the metals appear below is called the reactivity series

|  |  |  |
| --- | --- | --- |
| **reactive unreactive** | | |
| K Na Li Ca Mg Al Zn Fe H Cu Ag Au | | |
| Reacts with acid to produce salt and hydrogen | | no reaction with acid as metal  cannot displace hydrogen from the acids |
| reacts with cold water to produce metal hydroxide and hydrogen | reacts with steam to produce a metal hydroxide/oxide and hydrogen | no reaction with water as it cannot displace hydrogen |

**Displacement reactions**

When a metal is pure and unreacted it exists in the form of atoms. When a metal reacts with a non-metal, the metal forms positive ions. Therefore metals in solutions of ionic compounds (metal salts) are in the form of positive ions.

A displacement reaction is a reaction in which the atoms of a more reactive metal displace the ions of a less reactive metal from its compound to become atoms again. This is because the atoms of the more reactive metal have much greater tendency to form ions than atoms of the less reactive metal. The more reactive the metal, the easier it forms ions. The less reactive a metal, the greater its tendency to exist as atoms.

The table below shows two examples of displacement reactions.

|  |  |
| --- | --- |
| Reaction between magnesium atoms and copper sulphate (copper ions – metal salt) to form magnesium ions and copper atoms. Magnesium forms ions much easier than copper. As the copper ions are removed the solution loses its blue colour. **Magnesium displaces copper.** | Reaction between iron atoms and copper sulphate (copper ions) to form iron ions and copper atoms. Iron forms ions much easier than copper. The copper atoms are being deposited onto the iron nail. **Iron displaces copper.** |
|  |  |
| magnesium + copper sulphate  → magnesium sulphate + copper  Mg (s) + CuSO4 (aq) → MgSO4 (aq) + Cu (s) | iron + copper sulphate  → iron (II) sulphate + copper  Fe (s) + CuSO4 (aq) → FeSO4 (aq) + Cu (s) |

Exercises

1. Predict whether or not the following reactions will take place. For those that will take place write word equations and balanced symbol equations.

a) magnesium + copper (II) oxide

b) iron + aluminium oxide

c) calcium + magnesium oxide

1. Predict whether or not the following reactions will take place. For those that will take place write word equations and balanced symbol equations.

a) magnesium + calcium nitrate

b) iron + copper (II) chloride

c) copper + silver nitrate

1. The table below shows the results obtained from a displacement experiment. Place the metals

in order of increasing reactivity with the least reactive first

most reactive: …………….. ………………….. least reactive: ……………….

|  |  |  |  |
| --- | --- | --- | --- |
|  | salt solutions of metals (e.g. sulphates) | | |
| metal | solution of X | solution of Y | solution of Z |
| X | no reaction | reaction | reaction |
| Y | no reaction | no reaction | reaction |
| Z | no reaction | no reaction | no reaction |

**Reduction and oxidation - redox**

1. **Oxidation and reduction in terms of loss and gain of oxygen**

|  |
| --- |
| 2.18 know the conditions under which iron rusts  2.19 understand how the rusting of iron may be prevented by:  • barrier methods • galvanising  • sacrificial protection.   * 1. understand the terms: oxidation, reduction, redox, oxidising agent, reducing agent in terms of gain or loss of oxygen and loss or gain of electrons. |

|  |
| --- |
| Oxidation = a reaction during which a substance gains oxygen /gains mass |

#### Examples of oxidation reactions

1. Heating of copper to form copper oxide, which is black powder, the copper is oxidised and its mass increases as oxygen is added.

2Cu (s) + O2 (g) → 2 CuO (s)

1. Burning of magnesium to form magnesium oxide; the magnesium is oxidised and its mass increased.

2Mg (s) + O2 (g) → 2 MgO (s)

Other examples of oxidation reactions:

* rusting or corrosion of iron.
* combustion of fuels
* respiration

#### Examples of reduction reactions:

Heating of mercury oxide; the mercury oxide is reduced as it loses its oxygen (mass is decreased).

2HgO (s) → 2Hg (l) + O2 (g)

|  |
| --- |
| *Reduction = a reaction during which a substance loses oxygen (loses mass)* |

*A redox reaction = a reaction during which both an oxidation and reduction occur. Oxygen is lost by one reactant and taken by another reactant.*

Example of redox reaction: Reaction between copper oxide and hydrogen.

**reduction** (oxygen is being lost)

CuO (s) + H2 (g) Cu (s) + H2O (g)

**oxidising reducing**

**agent agent**

**oxidation** (oxygen is gained)

In this reaction:

* the copper oxide has been reduced by the hydrogen which has been oxidised;
* the CuO is the oxidising agent;
* the hydrogen is the reducing agent.

In any redox reaction, the reducing agent, which gains the oxygen, causes the reduction of the other substance and by doing so becomes oxidised. The oxidising agent causes the oxidation of the other substance but itself becomes reduced in the process.

Complete the following passage.

When during a reaction a substance, usually an element, gains oxygen we say that the substance has

been \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and an \_\_\_\_\_\_\_\_\_\_\_\_\_\_ has taken place.

When during a reaction, a compound loses oxygen we say the compound has been \_\_\_\_\_\_\_\_\_\_\_\_\_\_

and a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ has taken place.

During a reduction often the oxygen that is lost goes to another substance which then becomes

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ . A reaction during which one substance is oxidised and another substance is

reduced is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ reaction.

A substance that causes another substance to lose oxygen and becomes itself oxidised is called a

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

A substance that loses oxygen to another substance becomes \_\_\_\_\_\_\_\_\_\_\_\_\_ and causes the other

substance to be oxidised; such a substance is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ .

For each of the redox reactions below state which substance is the oxidised and which is reduced. Also identify the oxidising agent and the reducing agent for each reaction.

(a) aluminium + iron oxide ⎯→ iron + aluminium oxide

(b) magnesium + carbon dioxide ⎯→ magnesium oxide + carbon

(c ) tin oxide + carbon ⎯→ tin + carbon monoxide

(d) NO2 (g) + SO2 (g) ⎯→ NO (g) + SO3 (g)

(e) Fe (s) + CuO (s) ⎯→ FeO (s) + Cu (s)

(f) Fe2O3 (s) + 3CO (g) ⎯→ 2Fe (s) + 3CO2 (g)

**Oxidation and reduction in terms of loss and gain of electrons**

OILRIG = **O**xidation **I**s **L**oss (of electrons) **R**eduction **I**s **G**ain (of electrons)

**Redox reaction = a reaction during which an electron(s) is transferred from one substance to another.**

How do we know if a substance is oxidised or reduced? Check oxidation states.

How?

* The oxidation state of an atom =
* its ionic charge if it has formed an ion; or
* the number of electrons it shares with another atom of a different element in a covalent bond;
* the oxidation state of an element on its own is zero.
* The total of all oxidation states in a molecule is zero.
* if its oxidation state increases a substance has been oxidised
* if its oxidation state decreases, a substance has been reduced;
* if there are no changes in oxidation states then no oxidation or reduction has taken place.

Example of how oxidation states in a reaction are used to identify reducing and oxidising agent:

Reaction between copper oxide and hydrogen.

(copper has gained 2 electrons)

**reduction** (oxidation state of Cu has decreased)

CuO (s) + H2 (g) Cu (s) + H2O (g)

+2 -2 0 0 +1 -2

**oxidising reducing**

**agent agent**

**oxidation** (oxidation state has increased) hydrogen has lost electrons

Hydrogen as the reducing agent has reduced copper. Or copper oxide has oxidized hydrogen and is the oxidizing agent.

For each of the reactions below decide if it is a redox reaction or not; if it is a

redox reaction say which reactant is the reducing agent and which one is the oxidising agent:

(1) O2(g) + 2H2 (g) ⎯→ 2H2O (g)

(2) 2Fe(s) + 3Cl2 (g) ⎯→ 2FeCl3 (s)

(3) MgO(s) + 2HCl (aq) ⎯→ MgCl2 (aq) + H2O (l)

(4) Mg (s) + CuO (s) ⎯→ MgO (s) + Cu (s)

(5) Ca(OH)2 (s) + H2SO4 (aq) ⎯→ CaSO4 (s) + 2H2O (l)

(6) CuSO4.5H2O (s) ⎯→ CuSO4 (s) + 5H2O (l)

(7) Fe2O3 (s) + 3H2 (g) ⎯→ 2Fe (s) + 3H2O (g)

(8) 2NaBr (aq) + Cl2 (aq) ⎯→ 2NaCl (aq) + Br2 (aq)

(9) CaCO3 (s) ⎯→ CaO (s) + CO2 (g)

(10) 2Mg(s) + SO2 (g) ⎯→ 2MgO (s) + S (s)

(11) N2(g) + 3H2 (g) ⎯→ 2NH3 (g)

(12) 2AgNO3 (aq) + MgCl2 (aq) ⎯→ 2AgCl (s) + Mg(NO3 )2 (aq)

(13) CaCO3 (s) + 2HNO3 (aq) ⎯→ CO2 (g) + Ca(NO3)2 (aq) + H2O (l)

(14) C(s) + 2H2SO4 (aq) ⎯→ CO2 (g) + SO2 (s) + 2H2O (l)

The displacement reactions of the halogens that you studied in subsection 2b are also examples of redox reactions as the examples below show**.**

|  |
| --- |
| Cl2 (aq) + 2NaBr (aq) → Br2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces the bromine. The chlorine atoms are reduced (each gains an electron) and bromide ions are oxidized (each lose an electron). This displacement reaction is also a redox reaction:  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2Br- → Br2 +2e-  Ionic equation: Cl2 (aq) + 2Br- (aq) → Br2 (aq) + 2Cl- (aq) |
| Cl2 (aq) + 2NaI (aq) → I2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each loses an electron).  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2I- → I2 +2e-  Ionic equation: Cl2 (aq) + 2I- (aq) → I2 (aq) + 2Cl- (aq) |
| Br2 (aq) + 2NaI (aq) → I2 (aq) + 2NaBr (aq)  During this reaction, the bromine displaces the iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each lose an electron).  Half equations: reduction: Br2 + 2e- → 2Br- oxidation: 2I- → I2 +2e-  Ionic equation: Br2 (aq) + 2I- (aq) → I2 (aq) + 2Br- (aq) |

**Rusting**

Rusting is a reaction between iron, oxygen and water producing a brittle red-brown product which is iron (III) oxide which weakens any iron structure which is why this process is an expensive nuisance.

iron + oxygen + water → hydrated iron (III) oxide (=rust) (Fe2O3.xH2O)

The conditions necessary for rusting can be investigated using simple experiments as shown on the following site: <http://www.bbc.co.uk/scotland/learning/bitesize/standard/chemistry/metals/corrosion_rev1.shtml>

The conditions for rusting are: oxygen and water whilst any electrolyte e.g. sodium chloride speeds up the process.

The prevention of rusting should focus on ensuring all three reactants (iron, oxygen and water) do not come in contact with each other.

|  |  |  |  |
| --- | --- | --- | --- |
| method | how it works | method | examples |
| Painting the iron surface, applying oil and grease, covering in plastic | A barrier between iron and oxygen/water. | Cover the iron. However, when barrier is broken, rusting occurs | Bicycle chains (oil), bicycle brake cables (plastic), bridges (paint) |
| Galvanising (= covering steel with zinc) (still works even if zinc is scratched away at some places) | Zinc is more reactive/better at losing electrons than iron so the zinc is oxidized by the oxygen and water instead of the iron | Zinc can be scratched and iron will still not rust.  Need to keep checking if there is enough zinc left on the iron. | Watering cans |
| Sacrificial protection with a more reactive metal (e.g. zinc, aluminium or magnesium) (zinc is used in ships) (galvanizing is also a form of sacrificial protection) | More reactive metal is oxidized by the oxygen and water instead of the iron. | More reactive metal can be scratched and iron will still not rust.  Need to keep checking if there is enough zinc left on the iron. | Ships, underground pipes |

**Section 2e: extraction and uses of metals**

|  |
| --- |
| 2.22C know that most metals are extracted from ores found in the Earth’s crust and that unreactive  metals are often found as the uncombined element  2.23C explain how the method of extraction of a metal is related to its position in the reactivity series,  illustrated by carbon extraction for iron and electrolysis for aluminium  2.24C be able to comment on a metal extraction process, given appropriate information  *detailed knowledge of the processes used in the extraction of a specific metal is not required* |

The ease of obtaining a pure metal from its ore (rock that contains a useful element or compound) depends on the reactivity of the metal.

The less reactive metals are easily removed from their ores as it only takes heating with carbon or carbon monoxide to do this. Reactive metals are difficult to obtain as their oxides or salts are difficult to decompose so a lot of energy is needed which is why they can only be extracted by using electrolysis. The unreactive metals do not need extraction as they are found pure in nature although copper is electrolysed further as it needs to be very pure to be used as a conductor of electricity.

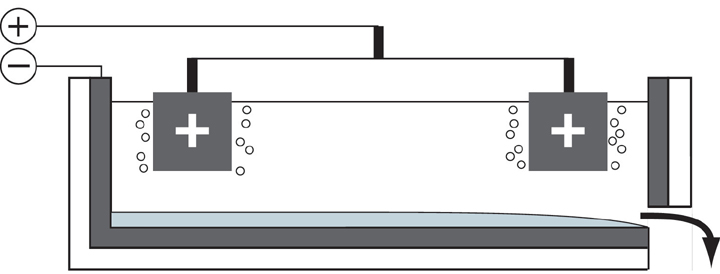
|  |  |
| --- | --- |
| **reactivity series** | **method of extraction** |
| potassium  sodium  calcium  magnesium  aluminium | reduction from ore using electrolysis |
| zinc  iron  lead  copper | reduction from ore by heating with carbon or carbon monoxide |
| silver  gold | found as uncombined element |

**Extraction of aluminium**

* Ore = bauxite; bauxite is a type of rock which has mainly aluminium oxide but also has some impurities like iron oxide and silica;
* starting chemical: aluminium oxide
* conditions:
* electrolytic cell with graphite electrodes; negative electrode in sides of the cell.
* aluminium oxide is molten by electrical heating; the large current which is used to electrolyse the aluminium oxide releases heat energy to keep the aluminium oxide molten.
* to help this process, cryolite is added to lower the operating temperature and the melting point of the mixture. The molten cryolite also acts as a solvent i.e. the aluminium oxide is dissolved in it.
* very large amounts of current which is why most aluminium smelters are near power stations and why this is an expensive process;
* after some time the positive graphite electrodes need to be replaced as they react with oxygen producing carbon dioxide.

Electrolytic cell used to electrolyse aluminium oxide:

Label all the parts and add the equations of all the reactions occurring in the electrolytic cell



Products:

|  |  |
| --- | --- |
| **anode** | **cathode** |
| Oxygen and carbon dioxide (formed when the oxygen reacts with the graphite rods; a reaction made easy by the large amount of heat energy; carbon electrodes get eroded away and need to be replaced every so often) | Molten aluminium |
| 2O2- (l) ⎯→ O2 (g) + 4e- | Al3+(l) + 3e- → Al (s) |
| overall equation : 4Al3+(l) + 12e- ⎯→ 4Al (s)  6O2- (l) ⎯→ 3O2 (g) + 12e-  4Al3+(l) + 6O2- (l) ⎯→ 4Al (s) + 3O2 (g)  or 2Al2O3 ⎯→ 4Al (s) + 3O2 (g) | |

**Extraction of iron**

Iron extraction is carried out in a blast furnace using the following raw materials:

|  |  |
| --- | --- |
| **raw material** | **what it is used for?** |
| coke | As fuel and to provide the carbon to make the reducing agent carbon monoxide |
| haematite | Contains the iron oxide |
| limestone | To remove impurities |
| hot air | Provides the heat and oxygen to make the reducing agent |

On the diagram of the blast furnace below label all chemicals.



The table summarises the five reactions which go on in the blast furnace.

|  |  |
| --- | --- |
| **Reactions** | **Equation** |
| Carbon burning to make carbon dioxide; this is important because it provides the heat for the reduction. | C (s) + O2 (g) ⎯→ CO2 (g) |
| Carbon dioxide reacting with carbon (at the top of the furnace there is little oxygen) to form carbon monoxide | C (s) + CO2 (g) ⎯→ 2CO (g) |
| **Carbon monoxide reduces the iron oxide**  **(this is the main reaction)** | Fe2O3 (s) + 3CO (g)⎯→ 2Fe (s) + 3CO2 (g) |
| Calcium carbonate decomposing | CaCO3 (s) ⎯→ CaO(s) + CO2 (g) (waste gas) |
| Calcium oxide reacting with silicon dioxide to form slag | CaO (s) + SiO2 (s) ⎯→ CaSiO3 (s) |

**Alloys**

|  |
| --- |
| 2.26C know that an alloy is a mixture of a metal and one or more elements, usually other metals or  carbon  2.27C explain why alloys are harder than pure metals |

* Sometimes we want to modify or change the properties of a metal so that it becomes more suitable for our its use e.g. it needs to be harder or stronger or have increased resistance to corrosion or increased heat or electrical resistance.
* We can improve the properties of metals by adding **small amounts** of another metal or non-metal. **An alloy is a mixture of a main metal with small amounts of one or more other elements usually other metals.**
* Alloys are harder (e.g. brass is harder than copper) or stronger than the pure metal. This is because the atoms in an alloy are of different sizes and they are arranged in such a way that the layers of atoms in the alloy do not slide easily over each other allowing the alloy to maintain its shape better.

Complete the particle diagram below showing an alloy of 2 different metals.

Arrangement of metal atoms in an alloy. Also add delocalised electrons!!!!

Why is it more difficult for the layers to slide past each other?

* The properties of alloys are different from the pure metal they are made from; their properties are easier to control and can be changed to ‘fit the purpose’.
* The components of an alloy are mixed when molten and the mixture is then allowed to cool.

Examples of alloys and their compositions and properties:

|  |  |  |
| --- | --- | --- |
| **alloy** | **composition** | **properties** |
| bronze | copper 90% and tin 10% | harder than pure copper |
| brass | copper 70% and zinc 30% | harder than pure copper and looks more like gold |
| solder | tin 50% and lead 50% | lower melting point than tin and lead |

**Uses of metals**

|  |
| --- |
| 2.25C explain the uses of aluminium, copper, iron and steel in terms of their properties  *the types of steel will be limited to low-carbon (mild), high-carbon and stainless* |

***Iron and steel***

The pig iron obtained from the blast furnace is still impure as there is about 5% carbon, silicon, phosphorus and sulfur in it which makes it rather hard and brittle and not very useful.

Its properties are modified by first lowering the amount of carbon, silicon, phosphorus and sulfur in it and then alloying it with other elements. When that happens the product is called steel.

There are hundreds of alloys of steel, each one with its own properties (some only differ in the carbon content).

Conversion to steel (=alloy); the properties of iron are changed by the controlled addition of other elements like carbon, nickel, chromium, tungsten, etc.

|  |  |  |  |
| --- | --- | --- | --- |
| **Type of steel** | **Typical composition** | **Properties** | **Uses** |
| Low-carbon (mild steel) | 99.75% iron, 0.25 % carbon | Easily worked, strong as it has lost most brittleness | Car bodies, ships, wires, nails, and machinery |
| Hard steel | upto 1.5 % carbon | Tough and brittle | Cutting tools, masonry nails |
| Stainless steel | 74% iron, 18% chromium, 8% nickel | Tough, does not corrode | Chemical plant, cooking tools and cutlery |

Pure iron or wrought iron is soft and is easy to shape and is therefore used for more decorative purposes such as railings and gates.

***Aluminium***

As aluminium is malleable, is a good conductor of electricity, has a low density (NOT LIGHT!!!!) but it is still strong so it is used for the following:

* in the manufacture of aircraft, car bodies, step ladders, … because of its strength and low density;
* in food containers because of its low density and malleability but also because it does not react with food or drink.
* in overhead electricity cables because of its low density and good electrical conductivity.

Copper

* Used in electrical wiring as it is a very good electrical conductor.
* Used in cooking utensils as it is a good conductor of heat and has low reactivity.
* Water pipes: low reactivity and high malleability.

**Section 2f and g: acids, alkalis, titrations and salt preparations**

**Indicators and pH**

|  |
| --- |
| 2.28 describe the use of litmus, phenolphthalein and methyl orange to distinguish  between acidic and alkaline solutions  2.29 understand how the pH scale, from 0–14, can be used to classify solutions as strongly  acidic (0–3), weakly acidic (4–6), neutral (7), weakly alkaline (8–10) and strongly alkaline (11–14)  2.30 describe the use of universal indicator to measure the approximate pH value of a solution |

Acid-base indicators are chemicals that show a colour or no colour when added in an acid but show a different colour when added to an alkali

The table below shows the different indicators and their colours in acids and alkalis that you need to know

|  |  |  |
| --- | --- | --- |
| **indicator** | **colour in acidic solution** | **colour in alkali solution** |
| litmus | red | blue |
| phenolphthalein | colourless | pink |
| methyl orange | red | yellow |

The indicators above do not tell how acidic or how alkaline a solution is i.e. they don’t tell us the pH of a solution, for that we need universal indicator, which is a mixture of indicators, that tells us the approximate pH of a solution, which apart from telling us whether the solution is acidic or alkali or not, also tells how acidic or how alkaline the solution is.

The pH is a measure of the amount of hydrogen ions there are in a solution; the lower the pH number, the higher the amount of hydrogen ions in the solution.

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| pH number | | | | | | | | | | | | | |
| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 |
| strong acidic weakly acidic neutral weakly alkaline strongly alkaline  (many H+ ions) (very few H+ ions) | | | | | | | | | | | | | |

**Acids**

|  |
| --- |
| 2.31a know that acids in aqueous solutions are a source of hydrogen ions, H+  2.36a understand that an acid is a proton donor |

|  |
| --- |
| **Acids are substances which dissociate in water and release hydrogen ions, H+.**  In fact it is the hydrogen ion that gives the acids its acidic properties.  In most reactions in which an acid is involved the acid donates its hydrogen ion(s) to the other chemical it reacts with e.g. a base.  **Acids are therefore also known as proton donors as a hydrogen ion consists basically of a single proton.**  Examples of acids: HCl (aq), H2SO4 (aq), HNO3 (aq), H2CO3 (aq).  For example: HCl → H+ + Cl-  HNO3 → H+ + NO3-  H2SO4 → 2H+ + SO42- |

**Bases**

|  |
| --- |
| 2.36b understand that a base is a proton acceptor  2.38 know that metal oxides, metal hydroxides and ammonia can acts as bases and that alkalis are bases that are soluble in water  2.31b know that alkalis in an aqueous solutions as a source of hydroxide ions, OH¯ |

A base is any compound that can **accept a proton (proton acceptor)** from an acid.

Common bases are metal oxides, and metal hydroxides as both the oxide and hydroxide in these bases are negative ions that can accept protons from acids as shown. Also metal carbonates can accept protons from acids which is why they react with them.

Most metal oxides and metal hydroxides are insoluble; a few metal oxides and metal hydroxides are soluble like NaOH (aq), KOH (aq), NH4OH (aq) and these are called alkalis (An alkali is a soluble base).

**Neutralization**

|  |
| --- |
| 2.35 understand acids and bases in terms of proton transfer  2.32 know that alkalis can neutralize acids |

A reaction between an acid and an alkali is called a **neutralization** reaction and it involves the hydrogen ions, H+, of the acid being changed into water molecules by reacting for instance with hydroxide ions, OH-, of an alkali.

4.4 define acids as sources of hydrogen ions, H+, and alkalis as sources of hydroxide ions, OH¯

The equation shows how the hydroxide ion from an alkali can accept a proton from an acid to form water.

H+ (aq) + OH- (aq) ⎯→ H2O (l)

As a result of this reaction, the hydrogen ions are used up and this increases the pH to 7 which is neutral. The alkali has neutralized the acid

The full equation of such a neutralization reaction between and acid and an alkali could be:

HCl (aq) + NaOH (aq) ⎯→ NaCl (aq) + H2O (l)

Another example of a neutralization reaction involving an acid and a base includes:

2HCl (aq) + MgO (aq) ⎯→ MgCl2 (aq) + H2O (l)

The equation below shows how the oxide form the base (MgO is not soluble) can accept protons from an acid to form water: 2H+ (aq) + O2- (s) ⎯→ H2O (l)

**Reactions of acids**

|  |
| --- |
| 2.37 describe the reactions of hydrochloric acid, sulfuric acid and nitric acid with metals, bases and metal carbonates (excluding the reactions between nitric acid and metals) to form salts |

Acids carry out the following reactions.

|  |
| --- |
| acid + metal ⎯→ salt + hydrogen  During this type of reaction the metal displaces the hydrogen ion from the acid that is then liberated as hydrogen gas molecules; the metal combines with the non metal of the acid to form a salt.  Example: 2HCl (aq) + Mg (s) ⎯→ MgCl2 (aq) + H2 (g)  Complete the following equations:   1. HCl (aq) + Ca (s) ⎯→ 2. H2SO4 (aq) + Ca (s) ⎯→ 3. HCl (aq) + Fe (s) ⎯→ FeCl3 (aq) 4. H2SO4 (aq) + Zn (s) ⎯→ 5. HCl (aq) + Al (s) ⎯→ |
| acid + metal oxide ⎯→ salt + water  During this type of reaction the metal displaces the hydrogen ion which reacts with the oxide ion to form water; the metal combines with the non metal of the acid to form a salt – the metal ionises.  Example: 2HCl (aq) + MgO (s) ⎯→ MgCl2 (aq) + H2O (l)  Complete the following equations:   1. HCl (aq) + CuO (s) ⎯→ 2. H2SO4 (aq) + CaO (s) ⎯→ 3. HCl (aq) + ZnO (s) ⎯→ 4. H2SO4 (aq) + Fe2O3 (s) ⎯→ 5. HNO3 (aq) + Al2O3 (s) ⎯→ |
| acid + metal hydroxide (alkali) ⎯→ salt + water  During this type of reaction the metal displaces the hydrogen ion which reacts with the hydroxide ion to form water; the metal ion combines with the non metal ion of the acid to form a salt.  Example: HCl (aq) + NaOH (aq) ⎯→ NaCl (aq) + H2O (l)  Complete the following equations:   1. HNO3 (aq) + NaOH (aq) ⎯→ 2. H2SO4 (aq) + KOH (aq) ⎯→ 3. HCl (aq) + NH4OH (aq) ⎯→ 4. H2SO4 (aq) + NH4OH (aq) ⎯→ 5. HNO3 (aq) + Cu(OH)2 (aq) ⎯→ |
| acid + metal carbonate ⎯→ salt + water + carbon dioxide  During this type of reaction the metal displaces the hydrogen ion which reacts with the carbonate ion to form water and carbon dioxide; the metal combines with the non metal ion of the acid to form a salt.  Example: 2HCl (aq) + CaCO3 (s) ⎯→ CaCl2 (aq) + H2O (l) + CO2 (g)  Complete the following equations:   1. HNO3 (aq) + CaCO3 (s) ⎯→ 2. H2SO4 (aq) + CaCO3 (s) ⎯→ 3. HCl (aq) + Na2CO3 (aq) ⎯→ 4. HNO3 (aq) + MgCO3 (s) ⎯→ |

**Preparation of salts**

|  |
| --- |
| 2.34 know the general rules for predicting the solubility of ionic compounds in water:  2.39 describe an experiment to prepare a pure, dry sample of a soluble salts starting from an insoluble  reactant  **2.33c describe how to carry out an acid-alkali titrations**  **2.40C describe an experiment to prepare a pure, dry sample of a soluble salt, starting from an**  **acid and alkali**  **2.41C describe an experiment to prepare a pure, dry sample of an insoluble salt, starting from**  **two soluble reactants**  2.42 *practical: prepare a sample of pure, dry hydrated copper(II) sulfate crystals starting from copper(II)*  *oxide*  **2.43C *practical: prepare a sample of pure, dry lead(II) sulfate*** |

Salts are ionic compounds which consist of a positive ion (metal ion or ammonium ion, NH4+) and a negative ion (non-metal ion like chloride ion, Cl-, or non-metal ion radical like sulphate or nitrate).

Most salts are soluble in water except a few as shown in the table below.

|  |  |  |
| --- | --- | --- |
|  | **soluble** | insoluble |
| Salts | All **nitrates**  All **chlorides** *except….*  All **sulphate**s *except…..*  sodium and potassium carbonates  All **sodium**, **potassium** and **ammonium** salts  Sodium, potassium and calcium hydroxides are soluble  (calcium hydroxide slightly soluble) | ….silver chloride, lead chloride  ….barium sulphate, lead sulphate, calcium sulphate  ….all other **carbonates**  **common hydroxides** |

How do we now make **samples of pure dry crystals** of a salt. The flowchart below will help you to identify the correct method and describes each type of experiment.

Is salt soluble?

yes

no

Precipitation (using 2 soluble reactants)

* Make 2 solutions each one containing 1 ion of the salt to be made (use soluble salts!!!)
* Add the solutions
* **Filter off precipitate** which is the insoluble salt
* **Wash** the residue with distilled water to remove any of the other salt solutions
* **Dry** with filter paper or on warm gauze

base soluble?

yes

Obtaining dry salt from solution

* **Concentrate solution until it becomes saturated** which is done by driving off **most** of the water by heating it. Solution is saturated when crystals form on a glass rod dipped in the solution.
* Stop heating and **allow solution to cool** and salt crystals to form (solubility of salt decreases as temperature drops).
* **Filter** the crystals to remove any leftover soluition
* **Dry** crystals with filter paper or place on warm gauze.

no

Titration (acid-alkali titration)

* Add acid via burette to alkali`+ indicator until neutral (or until temperature stops rising)
* Measure volume of acid added; throw away

solution as it contains the indicator;

* Add same volume of acid to same amount of

alkali without indicator

* Obtain dry salt from solution

Add excess base (insoluble reactant

* Add excess base to acid until no more dissolves (warm acid if necessary to increase the rate)
* Filter off excess base
* Obtain salt from solution (see to the left)

For each of the following salts, select suitable chemicals and decide which method of salt preparation you will use to make the salt.

|  |  |  |
| --- | --- | --- |
| **salt you are going to make** | **starting chemicals you will use** | **method** |
| magnesium sulfate |  |  |
| sodium nitrate |  |  |
| calcium chloride |  |  |
| barium chloride |  |  |
| ammonium nitrate |  |  |
| copper sulphate |  |  |

Complete the table below about the making of insoluble salts.

|  |  |
| --- | --- |
| salt | starting chemicals |
| 1. copper carbonate |  |
| 2. silver chloride |  |
| 3. lead sulphate |  |
| 4. lead chloride |  |
| 5. barium sulphate |  |

For each of the above examples write a balanced symbol equation.

**Section 2h tests for ions and gases**

**1. Cations**

|  |
| --- |
| 2.45 describe how to carry out flame tests  2.46 know the colours formed in flame tests for these cations: Li+, Na+, K+, Ca2+ using flame tests  2.47 describe the tests for these cations:  ii. NH4+ using sodium hydroxide solution and identifying the ammonia evolved  iii. Cu2+, Fe2+ and Fe3+ using sodium hydroxide solution |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Flame tests | **cation** | **colour of flame** |  | **cation** | **colour of flame** |
|  | Li+ | red |  | K+ | lilac |
|  | Na+ | yellow |  | Ca2+ | orange red |
|  | Cu2+ | blue-green |  |  |  |

The flame test is carried out in the following manner: A platinum or nichrome wire is cleaned in concentrated hydrochloric acid and dipped into the flame until the flame is colourless.

It is then cleaned in the acid again and then dipped into the sample that needs to be tested (you don’t need to do this if the sample to be tested is already in a solution). The sample is then held in a non-luminous flame.

1. Why is the wire cleaned thoroughly?
2. Suggest a reason why platinum is the best material for the wire.
3. Why is a non-luminous flame needed?
4. Why is the platinum wire dipped into the acid again?

Using sodium hydroxide solution

|  |  |  |
| --- | --- | --- |
| **cation** | **test** | **test result** |
| NH4+ | Add aqueous sodium hydroxide and warm, test the gas given off with damp red litmus paper. | Red litmus turns blue indicating presence of ammonia |
| Cu2+ | Add aqueous sodium hydroxide | Blue precipitate, soluble in excess |
| Fe2+ | Green precipitate, soluble in excess |
| Fe3+ | Red-brown precipitate, soluble in excess |

Questions

1. For the Cu2+, Fe2+and Fe3+ tests, find the name of the precipitate by writing a balanced chemical equation,

including state symbols, for each test. Assume each salt is a sulphate. The starting chemicals for the first

equation have been given.

(a) CuSO4 (aq) + NaOH (aq) → + precipitate =

(b) precipitate =

(c) precipitate =

2. For each of the above full symbol equations also write an ionic equation – again with state symbols.

(a)

(b)

(c)

3. Why must the litmus paper in the ammonium test be made damp before using it?

4. Why does the sodium hydroxide test not work with the Li+ and K+?

**2. Anions**

|  |
| --- |
| 2.48 describe simple tests for the anions:  i Cl−, Br− and I−, using acidified silver nitrate solution  ii SO42−, using acidified barium chloride solution  iii CO32−, using dilute hydrochloric acid and identifying the carbon dioxide evolved |

|  |  |  |
| --- | --- | --- |
| **anion** | **test** | **test result** |
| Cl- | Acidify with dilute nitric acid, then add aqueous silver nitrate | white precipitate |
| Br- | Acidify with dilute nitric acid, then add aqueous silver nitrate | pale cream precipitate |
| I- | Acidify with dilute nitric acid, then add aqueous silver nitrate | yellow precipitate |
| SO42- | Acidify with hydrochloric acid, then add aqueous barium chloride | white precipitate |
| CO32- | Add dilute acid, test gas produced with lime water | effervescence, CO2 produced |

In the tests for the halides and sulphate ion an acid is also added this is because the acid will react with other substances other than the halide ions or sulphate ion that could also make precipitates (e.g. carbonate ion.

Questions

1. For the Cl-, Br -and I- tests, find the name of the precipitate by writing a chemical equation for each test,

including state symbols. Assume each salt is a sodium salt. The starting chemical for the first equation

has been given.

(a) NaCl (aq) + AgNO3 (aq) → + precipitate =

(b) precipitate =

(c) precipitate =

2. Write full ionic equations for each of the reactions above.

(a)

(b)

(c)

3. Name the insoluble solid formed during the sulphate ion test.

1. Why is hydrochloric acid added in the sulphate test and nitric acid in the halide tests?

**3. Gases**

|  |
| --- |
| 2.44 describe simple tests for the gases hydrogen, oxygen, carbon dioxide, ammonia, chlorine. |

|  |  |  |
| --- | --- | --- |
| gas | test | test result |
| hydrogen (H2) | lighted splint | squeaky ‘pop’ |
| oxygen (O2) | glowing splint | relights |
| carbon dioxide (CO2) | bubble through limewater | milky, white precipitate |
| ammonia (NH3) | damp red litmus paper | turns blue |
| chlorine (Cl2) | damp red/blue litmus paper | bleaches |